Chemistry 1

Introductory Chemistry 1

Prepared by Prof. Paul M. Shiundu
with Dahir Mohamed Yusuf

African Virtual university
Université Virtuelle Africaine
Universidade Virtual Africana

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TABLE OF CONTENTS

I. Chemistry 1: Introductory Chemistry 1 ____ 3

II. Prerequisite Course or Knowledge ____ 3

III. Time ____ 3

IV. Materials ____ 4

V. Module Rationale ____ 4

VI. Content ____ 4

  6.1 Overview ____ 4
  6.2 Outline ____ 5
  6.3 Graphic organizer ____ 6

VII. General Objective(s) ____ 7

VIII. Specific Learning Objectives ____ 7

IX Pre-assessment ____ 11

X. Key Concepts (Glossary) ____ 17

XI. Compulsory Readings ____ 18

XII. Compulsory Resources ____ 20

XIII. Useful Links ____ 23

XIV. Teaching and Learning Activities ____ 30

XV. Synthesis Of The Module ____ 86

XVI. Summative Evaluation ____ 87

XVII. References ____ 91

XVIII. Student Records ____ 92

XIX. File Structure ____ 93

I. Chemistry 1: Introductory Chemistry 1

by Dahir Mohamed Yusuf, Amoud University

II. Prerequisite Courses or Knowledge

Entry requirements for undergraduate university studies in science. Students are expected to have done chemistry as a subject in high school as well as Mathematics.

III. Time

A total of 120 hours is required to complete this module and the allocation is as follows:

Unit	Topic	Approximate no. of hours
1	Matter and measurements	30 hours
2	Atomic structure and periodicity	30 hours
3	Molecules and compounds	30 hours
4	Chemical reactions and stoichiometry	30 hours

IV. Material

To successfully complete the learning activities in this module you will require Internet-connectivity to enable you to access and /or use the following tools and resources:

CD-ROMs;
Internet-based:-
Computer aided instruction (CAI)
Computer-based assessment (CBA)
Multimedia delivery (including video conferencing)
e-Library and data base utilization
Integrated learning environment
Interactive discussions/chat sessions; and
Recommended textbooks and other reference materials

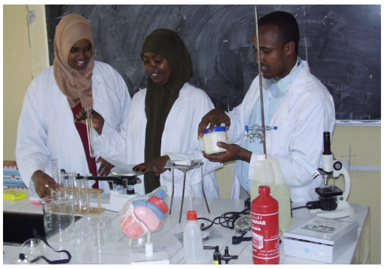
V. Module Rationale

This is an introductory module aimed at helping you to upgrade and consolidate your knowledge of some of the fundamentals of chemistry. To be successful, you must master concepts that form the building blocks of chemistry as well as acquire mathematical skills necessary to solve problems in this quantitative science. Chemistry, like driving a car, uses a language and skills of its own. Without a firm foundation in these fundamentals, a true understanding of chemistry is impossible. This module in part, begins to construct that foundation by introducing some key aspects of the chemist's view of matter. The module will enable the learner to learn some of the language of chemistry and develop an image of the physical world that will help you think like a Chemist.

VI. Content

6.1 Overview

This module covers introductory topics that are fundamental to understanding the subject of chemistry. In this module, we shall examine the structure of the smallest part of an element known as the atom (the building blocks of matter) and associated atomic models. The focus will be on the thought process involved in the development of the periodic table and its use in explaining the structure and properties of elements in "groups" and "periods". Both *microscopic* (in which matter is regarded as a collection of atoms and molecules) and *macroscopic* (associated with the bulk properties of matter) views of matter will be considered. The underlying principles that govern structures and shapes of simple molecules and ions will be covered as well as the nomenclature for writing formula for binary inorganic compounds, cations and anions. In addition, a review of the concepts behinds interpretation and balancing of chemical equations, stoichiometric calculations involving quantitative relations in chemical reactions, calculations of percent composition by mass and derivation of molecular formula from experimental data will also be made.



6.2 Outline

Unit I: Matter and measurements (30 Hours)

- Classifications of matter
- Physical and chemical properties of matter
- Matter and Energy
- Measurement

Unit II: Atomic Structure and Periodicity (30 Hours)

- The Atomic theory
- Development of the atomic theory and models
- Atomic Mass, Mass Number, Isotopes
- Writing Electronic Configurations
- The Periodic Table

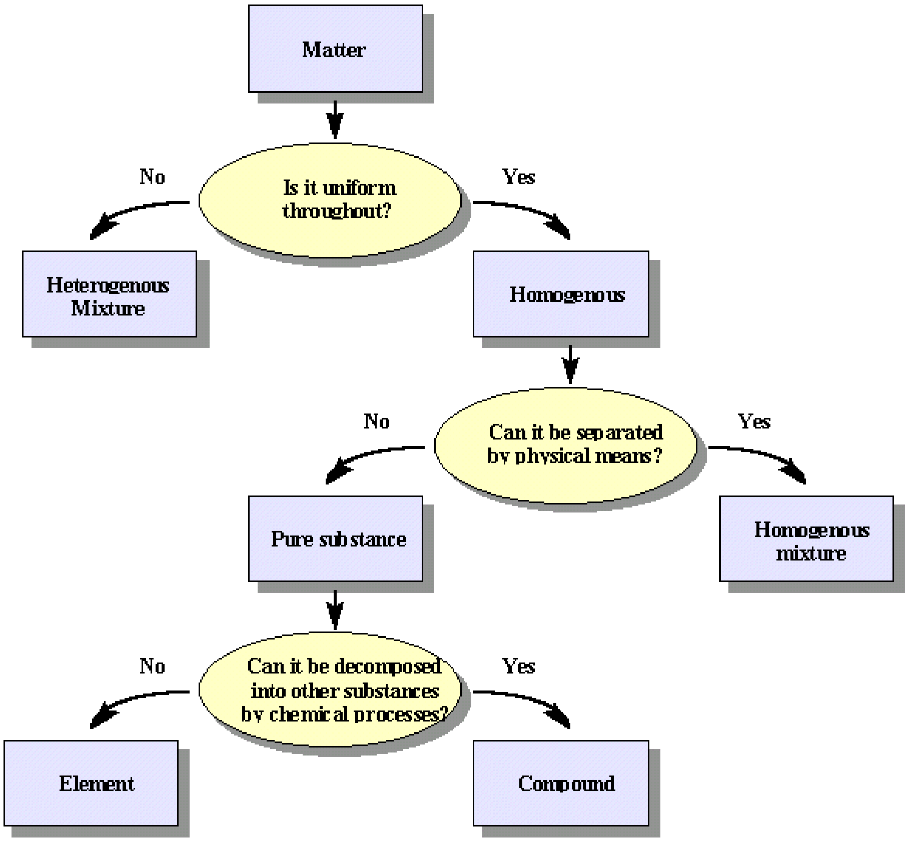
Unit III: Bonding, Molecules and Periodicity (30 Hours)

- The ionic and covalent bonds
- Intermolecular forces
- Formal charge and Lewis Structure
- Molecules and ions

Unit IV: Chemical Reactions and Stoichiometry (30 Hours)

- Percent composition of compounds
- Experimental determination of empirical and molecular formula
- Writing and balancing of chemical equations
- Stoichiometery and calculations
- Use of the mole concept in stoichiometric calculations

6.3 Graphic Organizer (can be drawn manually)


Matter
Is it uniform throughout?
No
Yes
Heterogenous Mixture
Homogenous
Can it be separated by physical means?
No
Yes
Pure substance
Homogenous mixture
Can it be decomposed into other substances by chemical processes?
No
Yes
Element
Compound

VII. General Objective(s)

The general objectives of this module are five-fold:

- To classify matter into types so as to make manageable the wealth of information about matter.
- To distinguish between matter and energy.
- To begin to classify elements in a systematic manner.
- To use the units of a measurement to do calculations involving that measurement.
- To use the laws of chemical combination to understanding chemical equations, balancing them and doing stoichiometric calculations.

VIII. Specific Learning Objectives

(Instructional Objectives)

Unit I: Matter and Measurement

At the end of the unit the student should be able to:

- Identify and explain physical and chemical properties
- Identify what distinguishes between solids, liquids and gases from each other
- Distinguish between matter and energy
- Use SI units to make scientific calculations and
- Determine uncertainties and error levels in chemical experiments

Unit II: Atomic Structure and Periodicity

At the end of this unit the learner should be able to:

- Discuss the development of the atomic theory and models
  and write electronic configurations of atoms.
- Understand the modern atomic theory and describe the structure of the atom.
- Write electronic configuration of atoms.
- Define atomic and mass number and carry out related calculations.
- Describe the thought process involved in the development of the periodic table.
- Use the periodic table to explain the structure and the properties of elements in groups and periods.

Unit III: Bonding, Molecules and Compounds

At the end of this unit the learner should be able to:

- Distinguish between ionic and covalent bonding.
- Describe and explain the criteria for determining types of intermolecular bonding.
- Use IUPAC nomenclature to name and write formula for binary inorganic compounds, cations and anions.
- Explain and predict chemical bonding, structure and geometrics of simple molecules and ions.
- Describe and apply the Lewis theory in making Lewis structure of binary compounds.

Unit IV: Chemical Reactions and Stoichiometry

At the end of this unit the learner should be able to:

- Calculate percent composition by mass and derive chemical formulae from experimental data.
- Balance and interpret chemical equations in terms of reactions and products.
- Define and use the idea of the mole to perform stoichiometric calculations involving quantitative relations in chemical reactions.

Unit Number	Learning Objective(s)
UNIT I **Matter and Measurement**	- Identify and explain the differences between solids, liquids and gases. - Identify and explain differences between physical and chemical properties of materials. - Distinguish between physical and chemical changes. - Discuss the differences between matter and energy. - Use the SI units of measurements to carry out relevant calculations associated with that particular measurement. - Determine uncertainties and error levels in chemical experiments.

UNIT II **Atomic Structure and Periodicity**	- Explain and distinguish between the various atomic theories. - Explain the modern atomic theory and structure and write electronic configurations of atoms. - Write the nuclear symbol of an element. - Define atomic mass, mass number and isotopes and carry out related calculations. - Explain and discuss the thought process involved in the development of the periodic table. - Explain the structure and properties of elements in groups and periods using the periodic table.
UNIT III **Bonding, Molecules and Compounds**	- Describe the two fundamental types of chemical bonding (i.e., ionic and covalent). - Know and compare properties of ionic compounds and covalent compounds. - Know and use the IUPAC chemical nomenclature for the systematic naming of chemical compounds. - Classify compounds as ionic or covalent. - Define and distinguish between empirical formulas, molecular formulas, and structural formulas for compounds. - Define, calculate and relate formula weights and molecular weights. - Use the IUPAC chemical nomenclature to name and write chemical formulas for binary inorganic compounds (covalent or ionic), simple ionic compounds (cations and anions).

UNIT IV
Chemical Reactions and Stoichiometry

- Explain the mole concept, and convert between grams, moles, and atoms and molecules.
- Determine mass percent composition of a sample from experimental data.
- Understand the concept of basic stoichiometry and related calculations.
- Determine mass percent composition of a compound from its formula.
- Determine empirical formula of a compound from its mass percent composition.
- Write balanced chemical equations from descriptions of chemical changes.
- Interpret chemical equations in terms of reactants and products.
- Classify chemical reactions as say, precipitation, neutralization, combustion, decomposition, etc.
- Write formation and combustion reactions for given compounds.

IX. Pre-assessment

Title of Pre-assessment :

Rationale: An atom is the smallest part of an element and constitute the building blocks of matter as we know it since matter is regarded as a collection of atoms and molecules. As the most basic unit that determines the structure and properties of elements, the learning of chemistry as a subject begins with atoms. For this reason, a set of pre-assessment questions are provided below which are meant to help you assess your previous understanding of this basic unit of matter and associated concepts. You will find some questions rather unfamiliar but these are meant to give you some idea of what to expect in this module.

Questions

1. An atom is

   a) The smallest particle in the nucleus
   b) The major component of an acid.
   c) The basic building block of matter
   d) A particle not larger than the proton.

2. The nucleus of an atom contains

   a) Clouds of gases and many other substances.
   b) Neutrons and protons.
   c) Electrons and protons.
   d) Protons and similar other particles.

3. If I meter =3.28 ft , then 50 meters will be

   a) 95 ft.
   b) 164 ft.
   c) 210.53 ft.
   d) 181.5 ft.

4. The mass of an atom is expressed in

   a) Pounds per volume.
   b) Atomic mass units.
   c) Volume per kilogram.
   d) None of the above.

5. Matter is anything that
   a) Occupies space and is solid only.
   b) Occupies a space only
   c) Occupies a space and has a mass.
   d) Occupies a space and is not liquid.

6. Mixtures are always
   a) Heterogeneous
   b) Heterogeneous in solution form
   c) inseparable
   d) Both heterogeneous and homogeneous.

7. Compounds can be
   a) Decomposed by chemical changes.
   b) Heterogeneous substances
   c) Only solid substances.
   d) Both solid and liquid substances.

8. Electrons are the
   a) Only particles seen in the nucleous of atoms.
   b) The particles that determine the mass of an atom
   c) Particles that are found in the shells of the atom.
   d) None of the above.

9. The atomic number
   a) Determines the type of atom.
   b) Is similar to the mass number.
   c) Is equal to neutrons plus protons.
   d) Is greater than the mass number.

10. The bond between atoms that share electrons
    a) Is called sigma bond.
    b) Is called a covalent bond.
    c) Is an ionic bond.
    d) Is both ionic and covalent bond.

11. A single bond results when
    - a) Two pairs of electrons are shared between two atoms.
    - b) Two or more pairs of electrons are shared.
    - c) Two ionic bonds are joined together.
    - d) Two electrons are shared between two atoms.

12. The measure of the ability of an atom to attract a pair of electrons
    - a) Is called simple attraction.
    - b) Is called electro negativity.
    - c) Is called ionic contraction.
    - d) All are correct.

13. Ionic bonds are often
    - a) Stronger than normal bonds.
    - b) Weaker than normal bonds.
    - c) Stronger than covalent bonds.
    - d) Weaker than covalent bonds.

14. If sodium hydroxide is added to hydrochloric acid
    - a) The product is chlorine gas.
    - b) The product is chlorine and hydrogen gases.
    - c) The product would be sodium chloride only.
    - d) The product would be sodium chloride and water.

15. An orbital is characterized by
    - a) The principle quantum number.
    - b) The azimuthal quantum number.
    - c) The magnetic quantum number.
    - d) a, b, and c

16. Elements in which the s and p orbitals in the outer shell are completely filled are:
    - a) All solid elements.
    - b) Both liquid and solid elements.
    - c) Solid, liquid, and gas
    - d) Noble gases.

17. A sodium atom can easily form an ionic bond with
    a) A carbon atom.
    b) A neon atom.
    c) A lead atom.
    d) A chlorine atom.

18. Hydrogen readily forms a covalent bond with
    a) Atoms of all elements.
    b) The halogens.
    c) Atoms in group III A.
    d) None of the above.

19. Gases are
    a) Substances that take the shape of their containers.
    b) Substances that take the volume of their containers.
    c) Not compressible.
    d) A and b

20. Select the correct symbol of calcium
    a) C
    b) Ca
    c) CA
    d) ca

Title of Pre-assessment : ______________________________

Answer Key

1. C
2. B
3. B
4. B
5. C
6. D
7. A
8. C
9. A
10. B
11. D
12. B
13. C
14. D
15. E
16. D
17. D
18. B
19. D
20. B

Notes to the student:

If you get 10 items or more correct you can consider that you are doing fine. However, if you get less than 8 items correct then you need to put in extra effort to perform well in this module.

Pedagogical Comments For Learners

Important Tips:

- You will be able to appreciate why various substances behave the way they do by understanding the theory and structure of atoms. This is quite fundamental in conceptualizing the very existence of inter- and intra bonding between the components that make up an element and molecules, respectively.
- Bohr's theory is of fundamental importance in understanding the behavior of matter in general and need to be mastered properly. It forms the foundation of this module.

X. Key Concepts (Glossary)

A pure substance*:* A substance with a definite chemical composition.

Atom*:* the smallest particle of an element that retains the identify and properties of the element and can take part in a chemical change.

Atomic number *(symbol Z):* the number of protons in the nucleus of each atom.

Compound: a substance that is formed when two or more elements combine chemically to form a substance with completely different properties.

Electron Configurati*on*: the arrangement of electrons in the orbitals of an atom.

Element*:* a substance that cannot be separated chemically into any simpler substances, e.g., Oxygen O_2; chlorine, Cl_2; carbon, C; and copper metal, Cu.

Isotopes: different varieties of an element, identical in chemical properties but differing slightly in atomic weight, i.e., If two atoms of an element have the same number of protons, but different numbers of neutrons, they are called ***ISOTOPES***.

Mass number (symbol A)*:* the total number of protons and neutrons in the nucleus of each atom.

Mole of a substance*:* the amount of substance that contains as many elementary entities (e.g, electrons, atoms, molecules) as there are atoms in exactly 12 grams (0.012kg) the carbon-12 isotope.

Molecule: the smallest mass of an element or compound capable of existing alone and possessing the properties of that element or compound.

Periodic Law: states that the properties of the elements are periodic functions of their atomic numbers.

Periodic Table: the arrangement of atoms in increasing order of their atomic numbers that puts atoms with similar properties in vertical columns.

Chemical Reaction: A chemical change which may be simple combination as in the case of two elements.

XI. Compulsory Readings

Reading #1

Complete reference: See PDF file named “Bishop_Book_1_eBook.pdf”

Abstract : This is a 29-page Chapter of a book entitled “An Introduction to Chemistry”; an Openware Source material. The chapter comprises of 5 relevant subsections namely; What is Chemistry and What can Chemistry do for you?; Suggestions for studying chemistry; The scientific method; Measurement and Units; and Reporting values from measurements.

Rationale: This chapter is an excellent reading material for people wishing to develop interest in learning and teaching chemistry as a subject. The first three subsections provide general but important information for developing interest in the subject of chemistry. The latter two subsections are directly relevant to Unit 1 of this module, which deals with “Matter and Measurements”. Within this chapter, you will find information on the origin of International System of Measurements (SI Units) and its relevance. The chapter illustrates the use of the International System of Measurements base units and their abbreviations to describe length, mass, time, temperature, and volume. The end of this Chapter contains a glossary section from where important definitions can be found. In addition, practice questions that are relevant to the topic are also provided.

Reading #2

Complete reference: See PDF file named “Bishop_Book_2_eBook.pdf”

Abstract : This is a 35-page Chapter 2 of a book entitled “The Structure of Matter and the Chemical Elements”; an Openware Source material. The chapter is made of 5 subsections whose corresponding subtitles are; Solids, Liquids, and Gases; The Chemical Elements; The Periodic Table of the Elements; The Structure of the Elements; and Common Elements. All these subtopics are very relevant to some of the contents of this module.

Rationale: This chapter begins a journey that will lead you to understanding of chemistry. It provides you with the basic knowledge of the chemical principles and underlying facts related to chemistry. The chapter will enable the learner to learn some of the language of chemistry and develop an image of the physical world that will help the learner think like a Chemist.

Reading #3

Complete reference: See PDF file named “Bishop_Book_3_eBook.pdf”

Abstract : This is a 55-page Chapter 3 of a book entitled “Chemical Compounds”; an Openware Source material. The chapter comprises of 5 subsections whose corresponding subtitles are; Classification of Matter; Compounds and Chemical Bonds; Molecular Compounds; Naming Binary Covalent Compounds; and Ionic Compounds.

Rationale: Most of the substances that we see around us are consist of two or more elements that have combined chemically to form more complex substances called compounds. In this chapter, you will learn to (a) define the terms mixture and compound more precisely, (b) distinguish between the terms elements, compounds, and mixtures, (b) describe how elements combine to form compounds, (c) construct systematic names for some chemical compounds, and (d) describe the charateristics of certain kinds of chemical compounds. The chapter will also expand your ability to visualize the basic structures of matter.

Reading #4

Complete reference: See PDF file named “Bishop_Book_4_eBook.pdf”

Abstract : This is a 33-page Chapter 4 of a book entitled “An Introduction to Chemical Reactions”; an Openware Source material. The chapter comprises of 2 subsections whose corresponding subtitles are; Chemical Reactions and Chemical Equations; and Solubility of Ionic Compounds and Precipitation Reactions.

Rationale: Once one understands the basic structural differences between different kinds of substances, the next important step is to learn how chemical changes takes place as one subtsnce is converted into the other. Such chemical changes are typically the concern of a chemist. Chemists often want to know, if anything happens when one substance encounters another. Whether the substances change and if so, why and how? In this chapter, the learner will find out about a chemical change that occurs when, say a solid dissolves in water, etc. These kind of changes can be described with chemical equations. This chapter begins with a discussion of how to interprete and write chemical equations.

XII. Compulsory Resources

Resource #1

Complete reference: The relevant titles of the simulations are: Simulated hydrogen and helium atoms and Hydrogen Ion Simulation. The corresponding sites are:

http://www.visionlearning.com/library/x_linker.php?moid=2494 & http://www.visionlearning.com/library/x_linker.php?moid=2141 , respectively.

Abstract : Atoms are electrically neutral because the number of protons (+ charges) is equal to the number of electrons (- charges) and thus the two cancel out. Ions on the other hand are charged (+ or – charges). As the atom gets larger, the number of protons increases, and so does the number of electrons (in the neutral state of the atom). The first illustration linked above compares the two simplest atoms, hydrogen and helium. The second is an the animation that shows a positive hydrogen ion (which has lost an electron) and a negative hydrogen ion (which has gained an extra electron). The electron charge on the ion is always written as a superscript after the atom's symbol, as seen in the animation.

Rationale: We now know that atoms are extremely small. One hydrogen atom(the smallest atom known) is approximately 5 x 10^{-8} mm in diameter. To put that in perspective, it would take almost 20 million hydrogen atoms to make a line as long as this dash -. Most of the space taken up by an atom is actually empty because the electron spins at a very far distance from the nucleus. For example, if we were to draw a hydrogen atom to scale and used a 1-cm proton (about the size of this picture -), the atom's electron would spin at a distance of ~0.5 km from the nucleus. Now, this is generally, difficult to visualize and conceptualize the very existence of an atom with its subatomic particles. As much as the learner will be able to read material pertaining to the "atom", it is bound to be difficult to picture this "supposedly" small entity that all matter is made of. This simulated illustration, is meant to show the learner what an atom and an ion looks like, the positioning and movement of the electron(s) (where applicable) with respect to the nucleus and compares the two simplest atoms of hydrogen and helium. It is hoped that with these illustrations, the learner will be better placed in conceptualizing the existence of both an atom and an ion.

Resource #2

Complete reference: The relevant titles of the simulations are: Bohr's Atom: Quantum Behavior in Hydrogen. The corresponding site is:

Javascript:WinOpen('/library/flash_viewer.php?oid=1347&mid=51','VLFlash',770,660);

Abstract : Under Bohr's theory, an electron's energy levels (also called electron shells) can be imagined as concentric circles around the nucleus. Normally, electrons exist in the ground, meaning they occupy the lowest energy level possible (the electron shell closest to the nucleus). When an electron is excited by adding energy to an atom (for example, when it is heated), the electron will absorb energy, «jump» to a higher energy level, and spin in the higher energy level. After a short time, this electron will spontaneously «fall» back to a lower energy level, giving off a quantum of light energy. Key to Bohr's theory was the fact that the electron could only «jump» and «fall» to precise energy levels, thus emitting a limited spectrum of light. The animation linked above simulates this process in a hydrogen atom. Concept simulation - Reenacts electron's «jump» and «fall» to precise energy levels in a hydrogen atom. (Flash required)

Rationale: According to Bohr, the line spectra phenomenon showed that atoms could not emit energy continuously, but only in very precise quantities (he described the energy emitted as quantized). Because the emitted light was due to the movement of electrons, Bohr suggested that electrons could not move continuously in the atom (as Rutherford had suggested) but only in precise steps. Bohr hypothesized that electrons occupy specific energy levels. When an atom is excited, such as during heating, electrons can jump to higher levels. When the electrons fall back to lower energy levels, precise quanta of energy are released as specific wavelengths (lines) of light. This animation affords the learner an opportunity to see existence of discrete lines and the origin of light emission upon excitation of an electron.

Resource #3

Complete reference: The relevant titles of the simulations are: The Lithium atom and Atomic structure animation table. The latter provides further details through a table linked which shows the electron configurations of the first eleven elements. The corresponding sites are:

http://www.visionlearning.com/library/x_linker.php?moid=2495 & http://www.visionlearning.com/library/x_linker.php?moid=2496 , respectively.

Abstract: Not only did Bohr predict that electrons would occupy specific energy levels, he also predicted that those levels had limits to the number of electrons each could hold. Under Bohr's theory, the maximum capacity of the first (or innermost) electron shell is two electrons. For any element with more than two electrons, the extra electrons will reside in additional electron shells. For example, in the ground state configuration of lithium (which has three electrons) two electrons occupy the first shell and one electron occupies the second shell. This is illustrated in the animation linked below.

Rationale: The concept of filling electron shells with electrons is just as abstract as conceptualizing the existence of at atom. This animation for the lithium atom would help the learner appreciate these phenomena and is valuable to the learner.

XIII. Useful Links

Useful Link #1

Title: Matter: Atoms from Democritus to Dalton,» Visionlearning Vol. CHE-1 (1), 2003.

URL : http://www.visionlearning.com/library/module_viewer.php?mid=49

Screen capture :

Description: This is an Open Source Visionlearning Volume CHE-1(1) book material by Anthony Carpi, Ph.D. entitled "Matter: Atoms from Democritus to Dalton". Within this site, you will be linked to other useful sites.

Rationale: This is a useful link as it provides the learner with an important chronology of developments from early theories governing matter to the modern understanding of matter as we know it today. In here, the origin of the "Law of conservation of mass" is articulated. Also, through this site you will have access to "Dalton's Playhouse": an interactive, virtual set of experiments that allow you to recreate classic experiments from the 19th century.

Useful Link #2

Title : Virtual Chemistry Experiments and Exercises

URL : http://www.chm.davidson.edu/ChemistryApplets/index.html

Screen capture :

Description: This is an open source site that allows the learner to carry out virtual experiments and exercise relevant to some sections of this module. Among the significant topics include "Atomic Struture" and "Chemical Bonding".

Rationale: This is a must-see site that enables the learner to carry out virtual chemistry experiments and exercises. The topics of significance to this module includes "Atomic Structure" and "Chemical bonding". Under the former, the learner will exposed to virtual chemistry experiments and exercises that deal with atomic and hybrid orbitals. The latter will have material on hybrid and molecular orbitals. In both situations, the Virtual Reality Modeling Language (VRML) will be used to display three-dimensional images of molecular structure and orbitals. Through the choice of the relevant links, the learner will be able to see (a) the isosurfaces of various atomic orbitals and some hybrid orbitals and Isosurfaces, radial distribution plots, and electron density plots of various atomic orbitals displayed. Virtual experiments showing shapes and orientation of s, p, and d orbitals are displayed. The significance of orbital overlap in forming a chemical bond is explained. Sigma, pi, and delta bonding interactions are also explained. The overlap of two s orbitals is illustrated among other important aspects of these topics.

Useful Link #3

Title : Atomic Theory I: The Early Days,» Visionlearning Vol. CHE-1 (2), 2003.

URL : http://www.visionlearning.com/library/module_viewer.php?mid=50

Screen capture :

Description: This is an Opensource, 2003 Volume CHE-1(2) Visionlearning material that covers the topic of «Atomic Theory I : The Early Days». The site features illustrations through appropriate links such as Simulated hydrogen and helium atoms, which compares the two simplest atoms, hydrogen and helium. The site also provides links to other relevant sites or modules such as Atomic Theory II; Matter, etc.

Rationale: The learner should find this site quite useful. The site features quizes for registered members of Visionlearning, an alphabetical glossary of relevant scientific terms through Visionlearning Glossary link, and an interactive practice exercise. In addition, the learner has access to illustrations formatted for printing on transparencies or displaying with a projector.

Useful Link #4

Title: General Chemistry Online

URL: http://antoine.frostburg.edu/chem/senese/101/index.shtml

Screen capture :

Description: This is a General Chemistry Online resource material that links the learner to various study material needed for this module. The range of topics that are linkable to appropriate databases and sites include : Measurement ; Matter ; Atoms & ions ; Compounds ; Chemical change ; The mole ; Energy and change ; The quantum theory ; Electrons in atoms ; The periodic table ; and Chemical bonds among other topics that are not directly relevant for this module.

Rationale: This is an invaluable resource site that has links to as many sites as there are topics of relevance to this module. The sites are well organized with accompanying Glossaryand Tutorial sections, just to mention two, for quick reviews by the learner. It is site that is highly recommended to the learner as it would make it very easy for the learner to gather relevant material. A special feature of General Chemistry Online is the existence of a key on: common compounds, exam guide, features,glossary,construction kits,simulations,toolbox and tutorial index associated with every topic covered. This make the sites pleasant to visit. Through the Companion Notes feature, which is available to every link, the learner has access to material such as: learning objectives; lecture slides; lecture notes; relevant links; frequently asked questions; etc. This makes following the material much easier.

Useful Link #5

Title: Atomic Structure, Bonding and Periodicity : Creative Chemistry : AS Chemistry Module 1 Worksheets and Guides

URL:http://www.creative-chemistry.org.uk/alevel/module1/index.htm

Screen capture :

Description: This page provides question sheets and practical guides on atomic structure, bonding and periodicity. The worksheets are available in PDF format (Adobe Reader required). It is the first part of the Advanced level (A- Level) section on the Creative Chemistry site produced by Dr Nigel Saunders, Harrogate Granby High School, UK. It is an Opensource material.

Rationale: The content of this page should provide the learner with good practice questions and practical guides on the topics of atomic structure, bonding, and periodicity. The target group here is supposed to be high school learners but the material affords the first year learner the opportunity to measure himself/herself against the target group.

Useful Link #6

Title : Lecture Help Pages with Solutions [Chemistry]

URL : http://chemistry2.csudh.edu/newlechelp/lechelpcs.html

Screen capture :

Description: Professor George Wiger from California State University, USA, has developed these exercises and calculations. They cover: matter and measurement; atoms and elements; molecules, ions and compounds; chemical equations and stoichiometry; reactions in aqueous solutions; energy and chemical reactions; atomic structures; atomic electron configurations and chemical periodicity; gases; solutions and their behaviour; chemical kinetics; chemical equilibria; acids and bases; and electron transfer reactions.

Rationale: This makes for a very useful resource for learners of this module. This site links the learner to several exercises with solutions that are meant to test his/her understanding of the numerous concepts learnt throughout this module. The variety of the exercises makes this a worthwhile resource for the learner. The other added advantage is the extent of coverage which takes care of virtually all the topics or units covered in this module.

Useful Link #7

Title : Atoms and the periodic table.

URL : http://www.chem1.com/acad/webtext/atpt/

Screen capture :

Description: Atoms and the periodic table: Introduction to the quantum theory of the atom, atomic structure and the periodic table is A Chem1 Virtual Textbook chapter authored by Stephen K. Lower, formerly of Simon Fraser University, Canada. The Chem1 Virtual Textbook is a collection of reference textbook chapters and tutorial units providing in-depth coverage of topics in college-level General Chemistry. This document covers its topic at a level appropriate for first-year college chemistry.

Rationale: This is a valuable 49-page pdf document that can be downloaded from the site: http://www.chem1.com/acad/pdf/c1atoms.pdf. The document contains important material relevant to this module. Material contained include: Quanta-a new view of the world; The Bohr Model of the Atom; Electron configurations of the elements; and Chemical periodicity. The last topic deals more with the organization of the periodic table; the shell model of the atom; sizes of atoms and ions; as well as the periodic trends in the ion formation. The learner should find it a useful resource for a good fraction of the material needed for this module.

Useful Link #8

Title: Chemical Bonding

URL : http://www.chem1.com/acad/webtext/virtualtextbook.html

Screen capture :

Description: This document is part of the Chem1 Virtual Textbook 77-page chapter by Stephen K. Lower of Simon Fraser University, Canada that covers the topic “Chemical Bonding” (http://www.chem1.com/acad/webtext/virtualtextbook.html) and may be reproduced for non-commercial purposes only.

Rationale: Chemical bonding refers to the set of principles and theories that govern the structure and stability of chemical substances, and thus of the re-arrangements that occur during chemical reactions in which one substance is transformed into another. The study of chemical bonding is thus one of the fundamental pillars of modern chemical science. This book chapter provides good coverage of this very important topic that forms part of a unit in this module. The subtopics within the book include: Bonds and molecules; Observable properties of chemical bonds; Why do chemical bonds form; The shared-electron model of chemical bonding; Polar and nonpolar bonds; The shapes of molecules:- the VSEPR model; Hybrid orbitals:- the valence bond model; and the Molecular orbital model. As much as the content may appear rather loaded, it will afford the learner, the opportunity to read wider.

Useful Link #9

Title: Chemical Principles: the Quest for Insight- Second Edition

URL : http://www.whfreeman.com/chemicalprinciples/

Screen capture :

Description: This website is a book companion site for ‘Chemical Principles: a Quest for Insight’, second edition (published by W H Freeman). This site has been developed to serve as an additional free resource for students and instructors using the textbook. Many resources to be used with this site and associated links require Macromedia Shockwave Player (version 8.5 or above), Macromedia Flash Player (version 6.0 or above), Apple QuickTime (5.0 or above), and Adobe Acrobat (version 6 or above) plug-ins.

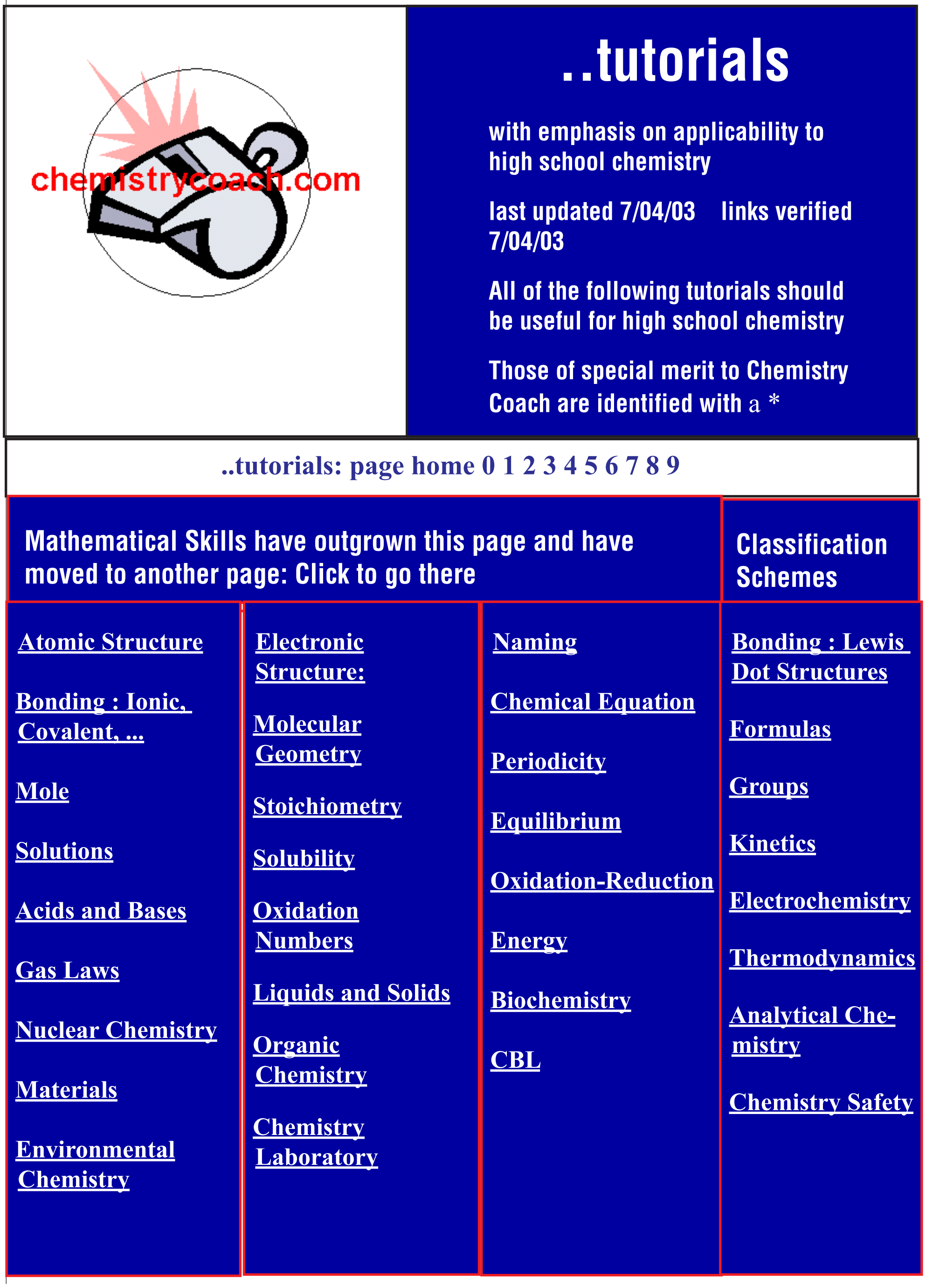
Rationale: Learners accessing this link plus associated links will find material contained here in very useful as they have a direct bearing on the module contents. The material includes chapter outlines, living graphs, animations, videos, molecular visualisations, simulations, exercises and links to related websites. Topics relevant to this particular module are: atoms (the quantum world), chemical bonds, molecular shape and structure, the properties of gases, liquids and solids, and, the elements (the first four and last four main groups), the d-block elements (metals in transition). Additional topics such as thermodynamics (first, second and third laws), physical equilibria, chemical equilibria, acids and bases, aqueous equilibria, electrochemistry, and chemical kinetics are beyond the scope of this module.

Useful Link #10

Title: Chemistrycoach.Com

URL : http://www.chemistrycoach.com/home.htm#Links

Screen capture :



..tutorials

with emphasis on applicability to high school chemistry

last updated 7/04/03 links verified 7/04/03

All of the following tutorials should be useful for high school chemistry

Those of special merit to Chemistry Coach are identified with a *

..tutorials: page home 0 1 2 3 4 5 6 7 8 9

Mathematical Skills have outgrown this page and have moved to another page: Click to go there

Classification Schemes

Atomic Structure

Bonding : Ionic, Covalent, ...

Mole

Solutions

Acids and Bases

Gas Laws

Nuclear Chemistry

Materials

Environmental Chemistry

Electronic Structure:

Molecular Geometry

Stoichiometry

Solubility

Oxidation Numbers

Liquids and Solids

Organic Chemistry

Chemistry Laboratory

Naming

Chemical Equation

Periodicity

Equilibrium

Oxidation-Reduction

Energy

Biochemistry

CBL

Bonding : Lewis Dot Structures

Formulas

Groups

Kinetics

Electrochemistry

Thermodynamics

Analytical Chemistry

Chemistry Safety

Description: This website links the learner to material content of virtually all the topics covered in this module. It is an openware source for material and the learner is encouraged to access it and its corresponding many links. The range of topics are clearly indicated in the screen capture above and each can be accessed individually by clicking onto them. The range of topics include: Atomic Structure, Electronic Structure, Naming, Bonding: Lewis Dot Structures; Bonding: Ionic, Covalent, …; Molecular Geometry; etc. Also included is a link to tutorials to assist the learner in understanding the material covered in this module.

Rationale: This is a must-access website that will lead the learner to virtually all the topics covered in this module. It is an openware resource that links the learner to the relevant material covered in this module. The learner is encouraged to access these links. Also included are links to tutorials on relevant topics covered in this module. All the tutorials available in this website should be very useful to material covered in high school chemistry and first year chemistry material.. Those of special merit to Chemistry Coach are identified with a *. The website has links to tutorials on virtually every topic covered in this module and should prove very useful to the learner. Tutorials relevant to each topic can be accessed via links associated with every topic. For instance, by clicking onto any of the following, the learner is linked to the relevant site that contains the relevant tutorial material: Original High School Tutorials; Relevant (High School) Chemistry Resources; Chemistry Coach.

XIV. Learning Activities

Learning activity # 1

Title of Learning Activity: Matter and Measurement

Specific Learning Objectives

- Define matter, and identify and explain the differences between the various states of matter (i.e., solids, liquids and gases).
- Identify and explain the differences between physical and chemical properties of matter.
- Define and distinguish between physical and chemical changes.
- Explain the differences between matter and energy.
- Understand measurements and apply the International standards of measurements to carry out relevant calculations associated with the particular measurement.
- Determine uncertainties and error levels in chemical experiments.

Summary of the learning activity

This activity comprises two interelated study topics: *Matter* and *measurement.* Under *Matter*, you will be introduced to the three forms of matter (gas, liquid, and solids) and the corresponding distinguishing characteristics between them at the molecular level. A classification of material properties as extensive, intensive, chemical, and physical properties will be made and their corresponding examples will be given. Under *Measurement*, you will be introduced to the International System of Measurements (SI units), which was developed to provide a very organized, precise and practical system of measurements that everyone in the world could use without ambiguity. You will be introduced to the construction of the SI units using seven (7) base units (such as metre, kilogram, second, kelvin, mole, ampere, and candela) from which other relevant units are derived. The last part of this activity centres on the need to report measurements in a way that not only shows the measurement's magnitude but also reflects its degree of uncertainty since all measurements are uncertain to some degree.

Key Concepts

Matter: anything that takes up space and has mass.

Space: that which is occupied by matter.

Mass: a measure of a body's resistance to a change in velocity.

Weight: refers to the force with which an object of a certain mass is attracted by gravity to the earth or to some other body that it may be near.

Physical change: any change NOT involving a change in the substance's chemical identity.

Chemical change: a process in which chemical bonds are broken and new ones are made.

Physical property: anything that can be observed without changing the identity of the substance.

Chemical property: characteristics which are exhibited as one substance is chemically transformed into another.

Law of conservation of mass: states that mass is neither created nor destroyed in a chemical reaction.

Law of definite proportions (also called law of definite composition):
states that in a pure chemical substance, the elements are always present in definite proportions by mass.

List of relevant readings

Bishop_Book_1_ebook.pdf
Bishop_Book_2_ebook.pdf
Bishop_Book_8_ebook.pdf
Physical and Chemical Properties.htm
States of Matter.htm
Physical and Chemical Changes.htm
Reactions.htm

List of relevant useful links

http://www.chemtutor.com/sta.htm
HTML site for material on “States of Matter”.
http://www.chemtutor.com/sta.htm
HTML site for material on “States of Matter”.
http://www.chemtutor.com/unit.htm
HTML site for material on “Units, Measures and Dimensions”.

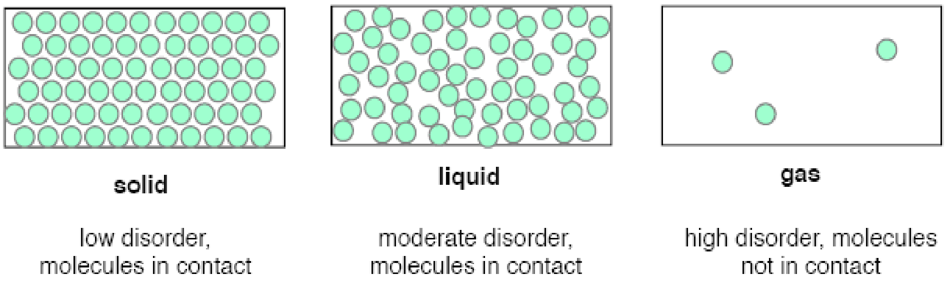
Detailed description of the activity

States of Matter:

Matter is defined as that which has **mass** and occupies **space**. Here, mass characterizes an object's resistance to change in its motion. **Space**, on the other hand is that which is occupied by matter.

There are three states of matter which exist in the environment of the Earth, namely: solid, liquid, and gas. These three states of matter can be distinguished from each other as follows:

Solid: have definite shape and definite volume;
Liquids: have indefinite shape and definite volume; and
Gases: have indefinite shape and indefinite volume.



solid
low disorder, molecules in contact

liquid
moderate disorder, molecules in contact

gas
high disorder, molecules not in contact

Figure 1: Schematic representations of the states of matter in terms of density and disorder.

By definition, **Indefinite shape** means that the sample in question takes on the shape of the container. For instance, if water is poured from a round container into a square container, the shape of the sample changes. On the other hand, **indefinite volume** means the sample would expand to fill the entire container. Only gases do this. Thus, if you had a gas in a 5-liter container and transferred the gas sample into a 10-liter container, the gas molecules would fill up the entire 10-liter volume. Note that in doing this, the gas becomes less dense.

Definite (for both shape and volume) means that the container makes no difference whatsoever. If 5-liters of liquid water is poured into a 10-liter container, the liquid would occupy 5-liters of the container and the other 5-liters would be empty. Suppose some water was frozen in the shape of a sphere and then put into a larger cubical shaped container. The spherical ball of ice would retain its spherical shape as well as retaining its volume even though it had been put in a container that was both larger and of a different shape.

Physical and Chemical Changes:

Physical Changes

A physical change is any change NOT involving a change in the substance's chemical identity. Here are some examples:

(1) Any phase change. For instance, moving between solid, liquid and gas involves only the amount of energy in the sample. There is no effect on the chemical identity of the substance. For example, water remains water, no matter if it solid, liquid or gas.

(2) Grinding something into a powder. Or the reverse process of making a bigger lump of stuff, say by melting lots of small pellets of copper into one big piece.

(3) Iron (and other metals) can be made to be magnetic. This change in no way affects the chemical identity of the element. Iron that is magnetized rusts just as easily as iron that is not magnetized.

The terminologies used to describe some of the commonly encountered physical changes are given below:

Change	**Name of change**
Solid to liquid	melting, fusion
Liquid to gas	boiling, evaporation
Solid to gas	sublimation
Gas to solid	deposition
Gas to liquid	condensation, liquefaction
Liquid to solid	freezing, solidification

An example of a sublimation process is conversion of dry ice from solid to gas. It is solid carbon dioxide and goes directly from the solid state to gas in the open atmosphere. One can make liquid carbon dioxide, but it must be done under extremely high pressures (e.g., 5 atmospheres).

Chemical Changes

A «chemical change» means that the reacting substances(s) are changed into new substances. The actual atoms involved remain, they are simply rearranged. The rearrangement is called a chemical reaction. For example:

$2H_2O \text{ ---> } 2H_2 + O_2$

is a chemical reaction in which water is broken down into the hydrogen and oxygen which make it up. Notice how the amounts of hydrogen atoms (four) and oxygen atoms (two) do not change from one side of the arrow to the other. However, the arrangements of the atoms is different. Some chemical bonds (the one involved in the water) have been broken and some new chemical bonds (the one in hydrogen and oxygen) have been formed.

A **chemical change** can also be defined as "a process in which chemical bonds are broken and new ones are made". Note that a process like grinding salt crystals into a fine powder does not involve the breaking of chemical bonds nor the formation of new ones, and so it is a **physical change**.

Physical and Chemical Properties

There are two types of property of matter. **Physical properties** describe the material as it is. **Chemical properties** describe how a material reacts, with what it reacts, the amount of heat it produces as it reacts, or any other measurable trait that has to do with the combining power of the material. Properties might describe a comparative trait (denser than gold) or a measured trait (17.7 g/cc), a relative trait (17.7 specific gravity), or an entire table of measurements in a table or graph form (the density of the material through a range of temperatures).

Physical Properties

Physical properties include such things as: color, brittleness, malleability, ductility, electrical conductivity, density, magnetism, hardness, atomic number, specific heat, heat of vaporization, heat of fusion, crystalline configuration, melting temperature, boiling temperature, heat conductivity, vapor pressure, or tendency to dissolve in various liquids. These are only a few of the possible measurable physical properties. **A physical property** of a pure substance can therefore be defined as anything that can be observed without changing the identity of the substance. The obervations usually consist of some type of numerical measurement, although sometimes there is a more qualitative (non-numerical) description of the property. There are many physical properties and here are some of the more common ones:

melting point	electrical conductivity	color	density
boiling point	thermal conductivity	odour	hardness
refractive index	atomic radius	ductility	

Note that, based on similarity in physical properties, elements or compounds can be grouped together. For instance, a group of substances can be called metals because they all possess a set of physical properties that are similar. For example, metals are ductile, malleable, conduct electricity and heat, and have luster. These all reflect a commonality of structure.

Note, however, that the similarities in a group do not extend to every property. For instance, while both tantalum and sodium are metals, tantalum's melting and boiling points are 2996 °C and 5425 °C, respectively and Sodium's is 98 °C and 883 °C, respectively. The wide disparity in the melting and boiling points between tantalum and sodium simply highlight the wide range that exists within the common structure that all metals have.

Chemical Properties

This one is more difficult. There really does not exist a set of properties that constitute chemical properties in the same way as physical properties. This is because chemical properties are tied to the change, whereas a given substance has a property (such as melting point) all to itself. Chemical properties include: whether a material will react with another material, the rate of reaction with that material, the amount of heat produced by the reaction with the material, at what temperature it will react, in what proportion it reacts, and the valence of elements.

Chemical properties can be defined as ***characteristics which are exhibited as one substance is chemically transformed into another***.

Here are some examples:

(1) iron rusting. When iron (an element, symbol = Fe) rusts, it combines in a complex fashion with oxygen to form a reddish-colored compound called ferric oxide (formula = Fe_2O_3). Not all substances rust.

(2) glucose, mixed with yeast, ferments to make alcohol. Glucose ($C_6H_{12}O_6$) is a chemical compound which enzymes in yeast can use to make ethyl alcohol (C_2H_5OH). Not all substances ferment.

(3) trinitrotoluene (TNT) reacts very, very fast when it is ignited. Among other products, it generates plenty of nitrogen gas and of heat. Inside an appropriate container, it can cause an explosion. Not all substances can make an explosion.

We can separate or purify materials based on the properties. For instance, we can separate wheat from chaff by throwing the mix into the wind. The less dense chaff is moved more by the wind than the denser wheat. We can separate a mixture of sand and iron filings by magnetism. The iron filings will stick to a magnet dragged through the mixture. We can separate ethyl alcohol (good old drinking alcohol) from water by boiling point. This process is called distillation. A mixture of water and insoluble material with alcohol mixed in it will release the alcohol as vapor at the boiling point of alcohol (78 °C). We can separate by solubility. A mixture of table salt and sand can be separated by adding water. The salt dissolves and the sand does not.

Exercise 1: What properties distinguish solids from liquids? Solids from gases? And liquids from gases?

Exercise 2: Describe a chemical change that illustrates the law of conservation of matter.

Exercise 3: Classify each of the following homogeneous materials as a solution, an element or a compound: Iron, water, carbon dioxide, oxygen, glucose , blood plasma.

Exercise 4: It is necessary to determine the density of a liquid to four significant figures. The volume of solution can be measured to the nearest 0.01cubic cm.

(a) What is the minimum volume of sample that can be used for the measurement? Ans. 10.00 cm^3.

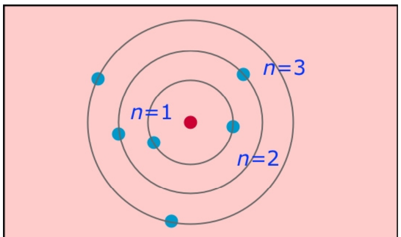
(b) Assuming the minimum volume sample determined in (a), how accurately must the sample be weighed (to the nearest 0.1g, 0.01g, ….), if the density of the solution is greater than 1.00g/ cm^3. Ans. Nearest 0.01g.

Exercise 5: Using exponential notations, express the following quantities in terms of SI base units:

(a) 0.13g (b) 5.23g

Exercise 6: How many significant figures are contained in each of the following numbers:

(a) 113 (b) 207.033 (c) 0.0820



In Bohr's model, electrons can only occupy orbits having integer values of *n*.

Learning activity # 2

Title of Learning Activity : Atomic Structure and Periodicity

Specific Learning Objectives

- Explain the development of the Atomic Theory
- Understand the modern atomic theory and describe the structure of the atom.
- Write electronic configurations of atoms.
- Define atomic and mass number and carry out related calculations.
- Describe the thought process involved in the development of the periodic table.
- Use the periodic table to explain the structure and the properties of elements in groups and periods.

Summary of the learning activity

A chemist's primary interest is the behaviour of matter. However, to understand the behaviour of matter we must first understand its internal structure. Prior to 1897, the internal structure of the atom had been a source of speculation for thousands of years. It is J.J. Thomson who discovered the electron, the first subatomic particle and was the first to attempt to incorporate the electron into a structure for the atom. Today, we know that atoms, which constitute elements, come together in different manners to form compounds and that the various manners by which atoms combine is attributable to their individual structures, which in turn influences the properties of the resultant compounds. It is therefore significant that we have a good knowledge of the structure of this smallest building block of an element. In this Unit, you will be expected to have a general knowledge of the various models that have been in existence inorder to appreciate the chronology of events leading to the present modern theory or model of the atom. The unit explains and describes the development of the atomic theory. The modern day theory about atoms is described and details about the structure of the atom is also included. A coverage on how to write electronic configuration of atoms; and definitions of the various terminologies governing atomic structure and the related calculations is made. The unit also covers the thought process involved in the development of the periodic table. The learner will be expected to use the periodic table to explain the structure and properties of elements in groups and periods.

Key Concepts

Atom: the smallest part of an element which can take part in a chemical change.

Atomic mass: the weighted average mass of all the atomic masses of the isotopes of that atom.

Atomic number: counts the number of protons in an atom.

Electron: sub-atomic particles which stream from the negative electrode when an electrical discharge passes through an attenuated gas.

Electron configurations: Lists the shells containing electrons written in order of increasing energy.

Element: a substance which it has not yet been possible to split up into a simpler substance, e.g., oxygen, chlorine, phosphorous, etc. It is the fundamental substance that cannot be broken down further by chemical means.

Isotopes: Atoms of the same element with the same number of protons, but different numbers of neutrons.

Law of multiple proportions: states that, suppose that we have samples of two different compounds formed by the same two elements. If the mass of one of the elements is the same in the two samples, the masses of the other element are in a ratio of small whole numbers.

Mass number: counts the number of protons and neutrons in an atom.

Periodic Law: States that all the elements in a group have the same electron configuration in their outermost shells.

Periodic chart: is a way to arrange the elements so as to show a large amount of information and organization.

Period: a line of elements as one reads across the periodic chart from right to left or vice-versa.

Group (or family): a line of elements in a periodic table as one reads the chart from top to bottom.

List of relevant readings

Atomic structure.htm
Atomic Theory I.htm
Atomic Theory II.htm
Notes Atoms & Ions.htm
Atoms & Elements.pdf
Atoms and Elements 2.pdf
Atoms and Isotopes.pdf
C1xatpt.pdf
Electron Arrangement and Periodic Law.pdf
Reviews of Elements, Compounds, and Mixtures.htm

List of relevant useful links :

http://www.chemtutor.com/struct.htm
HTML site for material on atomic structure.
http://www.chemtutor.com/elem.htm
HTML site for material on "elements".
http://www.chemtutor.com/perich.htm
HTML site for material on "Periodic Table".

Detailed description of the activity

Modern Atomic Theory

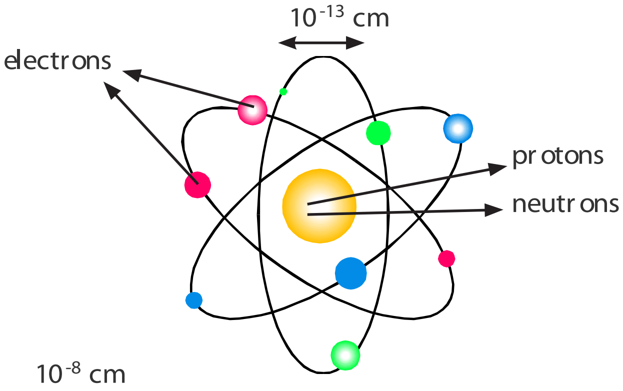
Elements are made up of atoms, the smallest particle that has any of the properties of the element. John Dalton, in 1803, proposed a modern theory of the atom based on the following assumptions:

1. Matter is made up of atoms that are indivisible and indestructible.
2. All atoms of an element are identical.
3. Atoms of different elements have different weights and different chemical properties.
4. Atoms of different elements combine in simple whole numbers to form compounds.
5. Atoms cannot be created or destroyed. When a compound decomposes, the atoms are recovered unchanged.

According to the modern atomic theory, **All** the matter around us is made of atoms, and all atoms are made of only three types of subatomic particles: protons, electrons, and neutrons. Furthermore, all protons are exactly the same, all neutrons are exactly the same, and all electrons are exactly the same. Protons and neutrons have almost exactly the same mass. Electrons have a mass that is about 1/1835 the mass of a proton. Electrons have a unit negative charge. Protons each have a positive charge. These charges are genuine electrical charges. Neutrons do not have any charge.

The subatomic particles and their corresponding charges (where applicable) are shown below:

Particle	Symbol	Charge	Relative Mass
Electron	e-	1-	0
Proton	p+	+	1
Neutron	n	0	1



Electron Configuration

Protons have a positive charge and electrons have a negative charge. Free (unattached) uncharged atoms have the same number of electrons as protons to be electrically neutral. The protons are in the nucleus and do not change or vary except in some nuclear reactions. The electrons are in discrete pathways or shells around the nucleus. There is a ranking or heirarchy of the shells, usually with the shells further from the nucleus having a higher energy as compared to those closer to the nucleus. As we consider the electron configuration of atoms, we will be describing the *ground state* position of the electrons. When electrons have higher energy, they may move up away from the nucleus into higher energy shells.

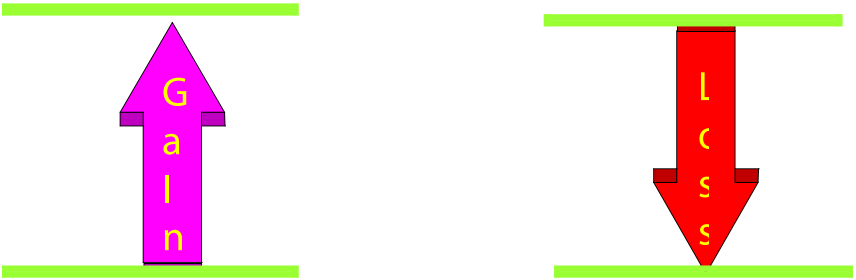
Characteristics of electrons

Electrons:

- Are extremely small mass.
- Are located outside the nucleus.
- Move at extremely high speeds in a sphere.
- Have specific energy levels

Electrons in atoms are arranged in discrete levels and when an electron absorbs energy (say, when an atom is heated), it is forced to "jump" to a higher energy level. This is often accompanied by emission of energy as the electron falls to a lower energy level.

Loss and Gain of Energy



Electron levels (or sometimes known as shells) contain electrons that are similar in energy and distance from the nucleus, with the low energy electrons being clossest to the nucleus. We identify the electron levels by numbers 1, 2, 3, 5, 6, ... The first shell (or energy level) 1 is lowest in energy, 2nd level is next and so on $1<2<3<4$ ….

The maximum number of electrons in any electron level is = $2n^2$, where ***n*** is the shell (or energy level) number. Therefore, for:

n = 1	$2(1)^2 = 2$
n = 2	$2(2)^2 = 8$
n = 3	$2(2)^3 = 16$

Order of filling electrons:

Recall that **ALL** electrons in the same energy level have similar energy and that

Shell 1	2 electrons
Shell 2	8 electrons
Shell 3	18 electrons (8 first, later 10).

Thus, the order of filling for the first 20 electrons can be represented as:

Shell	1	2	3	4
	2e	8e	8e	2e

Electron configuration lists the shells containing electrons and is written in order of increasing energy as shown below:

Element	**Shell**	**1**	**2**	**3**
He		2		
C		2	4	
F		2	7	
Ne		2	8	
Al		2	8	3
Cl		2	8	7

Periodic Law: States that all the elements in a group have the same electron configuration in their outermost shells.

Example: Group 2

Be	2, 2
Mg	2, 8, 2
Ca	2, 2, 8, 2

Orbital diagrams:

Orbital diagrams represent the filling and the arrangement of electrons within the various shells of an atom. Orbital diagrams:

- Use individual orbitals.
- Give subshell arrangement.
- Each orbital takes one electron before any other orbital in the same subshell can receive a second electron.

Characteristics of Subshells:

- Represent energy sublevels within an energy level.
- All electrons in a subshell have the same energy.
- The subshells are designated as *s, p, d, f* ..
- The sublevel energies are such that: s < p < d < f.

Corresponding to each energy level, there exists various sublevels as shown below:

Main Energy Levels	**Sublevels (or Subshells)**
n = 4	4s, 4p, 4d, 4f
n = 3	3s, 3p, 3d
n = 2	2s, 2p
n = 1	1s

Electrons allowed:

Since all electrons in the same sublevel have the same energy, **ALL** 2 s electrons have the same energy; ALL 2p electrons have the same energy which is slightly higher than the energy of the 2 s electrons. The total number of electrons allowed in each sublevel is as shown below:

s sublevel	2 electrons
p sublevel	6 electrons
d sublevel	10 electrons
f sublevel	14 electrons

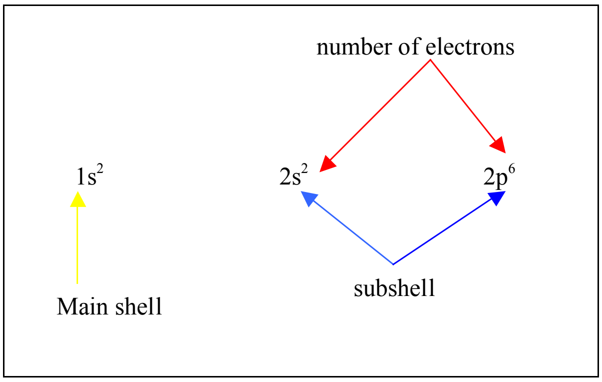
Electron Configuration:

This is written in order of increasing energy; with superscripts giving the number of electrons.

Example: Electron configuration of neon.

- The list of subshells containing electrons is as shown in the figure below:

$1\,s^2$ $2\,s^2$ $2\,p^6$

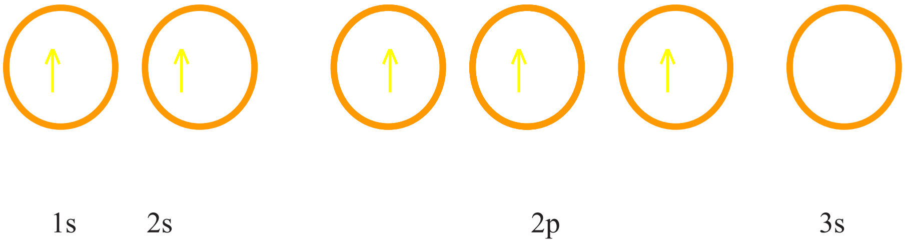


Writing electron configurations.

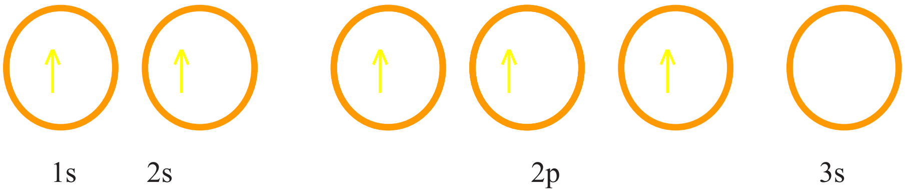
Below is a figure illustrating how to write the electron configuration of some elements.

H	$1s^1$				
He	$1s^2$				
Li	$1s^2$	$2s^1$			
C	$1s^2$	$2s^2$	$2p^2$		
S	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^4$

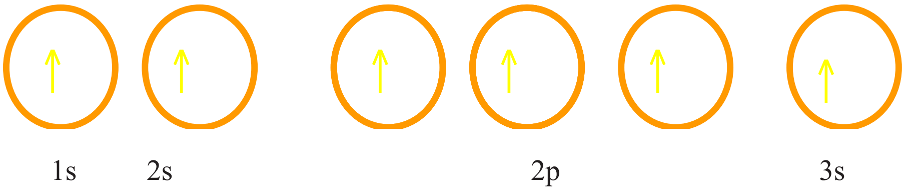
Example 1: The following represents the orbital diagram for a nitrogen atom:



Example 2: The orbital diagram for a fluorine atom is as below:



Example 3: The orbital diagram for a magnesium atom is:

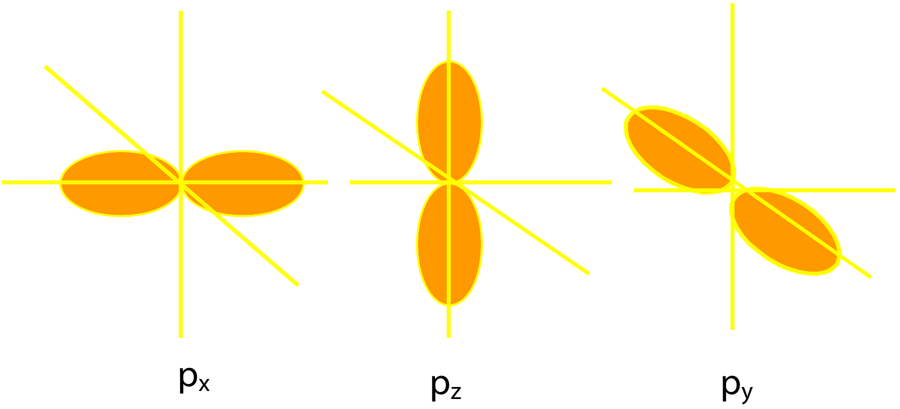


Problem 1: Write the orbital diagram for the electrons in an oxygen atom.

Problem 2: Write the orbital diagram for the electrons in an iron atom.

A diagramatic representation of the ***s*** orbitals is as below:

A diagramatic representation of the 3***p*** orbitals is as below:



In general, a hydrogen atom has only one proton and one electron. The electron of a hydrogen atom travels around the proton nucleus in a shell of a spherical shape. The two electrons of helium, element number two, are in the same spherical shape around the nucleus. The first shell only has one subshell, and that subshell has only one orbital, or pathway for electrons. Each orbital has a place for two electrons. The spherical shape of the lone orbital in the first energy level has given it the name '*s*' orbital. Helium is the last element in the first period. Being an inert element, it indicates that that shell is full. Shell number one has only one ***s*** subshell and all ***s*** subshells have only one orbital. Each orbital only has room for two electrons. So the first shell, called the **K shell**, has only two electrons.

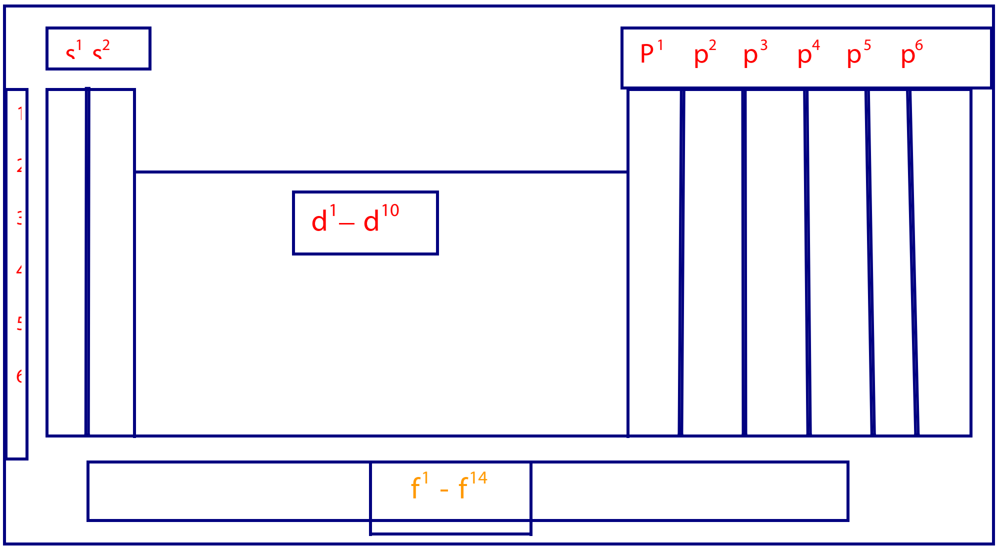
Beginning with lithium, the electrons do not have room in the first shell or energy level. Lithium has two electrons in the first shell and one electron in the next shell. The first shell fills first and the others more or less in order as the element size increases up the Periodic Chart, but the sequence is not immediately obvious. The second energy level has room for eight electrons. The second energy level has not only an *s* orbital, but also a *p* subshell with three orbitals. The *p* subshell can contain six electrons. The *p* subshell has a shape of three dumbbells at ninety degrees to each other, each dumbbell shape being one orbital. With the *s* and *p* subshells the second shell, the L shell, can hold a total of eight electrons. You can see this on the periodic chart. Lithium has one electron in the outside shell, the L shell. Beryllium has two electrons in the outside shell. The *s* subshell fills first, so all other electrons adding to this shell go into the *p* subshell. Boron has three outside electrons, carbon has four, nitrogen has five, oxygen has six, and fluorine has seven. Neon has a full shell of eight electrons in the outside shell, the L shell, meaning the neon is an inert element, the end of the period.

Beginning again at sodium with one electron in the outside shell, the M shell fills its *s* and *p* subshells with eight electrons. Argon, element eighteen, has two electrons in the K shell, eight in the L shell, and eight in the M shell. The fourth period begins again with potassium and calcium, but there is a difference here. After the addition of the **4***s* electrons and before the addition of the **4***p* electrons, the sequence goes back to the third energy level to insert electrons in a *d* shell.

The shells or energy levels are numbered or lettered, beginning with K. So K is one, L is two, M is three, N is four, O is five, P is six, and Q is seven. As the *s* shells can only have two electrons and the *p* shells can only have six electrons, the *d* shells can have only ten electrons and the *f* shells can have only fourteen electrons. The sequence of addition of the electrons as the atomic number increases is as follows with the first number being the shell number, the *s*, *p*, *d*, or *f* being the type of subshell, and the last number being the number of electrons in the subshell.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}\ 7p^6$

It is tempting to put an $\mathbf{8}s^2$ at the end of the sequence, but we have no evidence of an R shell. One way to know this sequence is to memorize it. There is a bit of a pattern in it. The next way to know this sequence is to SEE IT ON THE PERIODIC CHART. As you go from hydrogen down the chart, the Groups 1 and 2 represent the filling of an *s* subshell. The filling of a *p* subshell is shown in Groups 3 through 8. The filling of a *d* subshell is represented by the transition elements (ten elements), and the filling of an *f* subshell is shown in the lanthanide and actinide series (fourteen elements). See figure below.



Atoms and Elements

Elements are defined as the fundamental substances that cannot be broken down further by chemical means. Elements are the building blocks of the universe. They are pure substances that form the basis of all the materials around us. Some elements can be seen in pure form, such as mercury in a thermometer; some we see mainly in chemical combination with others, such as oxygen and hydrogen in water. We now know of approximately 116 different elements and each of the elements is given a name and a one- or two-letter abbreviation. Often this abbreviation is simply the first letter of the element; for example, hydrogen is abbreviated as H, and oxygen as O. Sometimes an element is given a two-letter abbreviation; for example, helium is He. When writing the abbreviation for an element, the first letter is always capitalized and the second letter (if there is one) is always lowercase.

A sample of an element contains only one kind of atom in the sample. Suppose you had a very very "pure" lump of copper in your hand. The ONLY type of atom in that lump is copper, i.e., in the lump there are trillions and trillions and trillions of copper atoms. NOTHING else. If one were to heat the lump of copper, it would melt and eventually vaporize. The smallest unit of the copper, called the atom, would remain unaffected by this. The atoms of copper would be in the solid state, the liquid state or the gaseous state, but they would be EXACTLY the same in each state. Thus, an atom could be viewed as the smallest subdivision of an element which still retains the properties of that element. In fact, a very good definition of an atom is: the smallest part of an element that can enter into a chemical combination.

An ***atom*** is the single unit of an element. It is the most basic unit of the matter that makes up everything in the world around us. Each atom retains all of the chemical and physical properties of its parent element. The atom however, has smaller particles called protons, neutrons, and electrons arranged in its structure. Atoms are electrically neutral because the number of protons (+ charges) is equal to the number of electrons (- charges) and thus the two cancel out. As the atom gets larger, the number of protons increases, and so does the number of electrons (in the neutral state of the atom).

All atoms of an **element** have the same number of protons. Atoms of different elements are distinguished from each other by their number of protons (the number of protons is constant for all atoms of a single element; the number of neutrons and electrons can vary under some circumstances). In order to identify this important characteristic of atoms, the term ***atomic number***, Z, is used to describe the number of protons in an atom. For instance, $Z = 1$ for hydrogen and $Z = 2$ for helium.

Another important characteristic of an atom is its weight or ***atomic mass***. The weight of an atom is roughly determined by the total number of ***protons*** and ***neutrons*** in the atom. While protons and neutrons are about the same size, the electron is more than 1,800 times smaller than the two. Thus the electron's weight is inconsequential in determining the weight of an atom. A count of the number of protons and neutrons in an atom is also referred to as the ***mass number***.

The number of neutrons in an atom can also vary. Atoms having the same number of protons, but different numbers of neutrons are known as ***isotopes***. Isotopes are essentially atoms of the same element (since they have the same atomic number) but having different mass numbers. For example, normally hydrogen contains no neutrons. However, an isotope of hydrogen does exist that contains one neutron (commonly called deuterium). Thus, the atomic number (Z) is the same in both isotopes; however the **atomic mass** increases by one in deuterium as the atom is made heavier by the extra neutron. Another example includes the isotopes of chlorine which are chlrorine - 35 and chlorine - 37.

The Periodic Chart of the Elements

The **Periodic Chart of the Elements** is primarily a way to arrange the elements so as to show a large amount of information and organization. As you read across the chart from right to left, a line of elements is a **Period** whereas as one reads down the chart from top to bottom, a line of elements is a **Group** or **Family**. The elements are numbered beginning with hydrogen, number one, in integers up to the largest number. The integer number in the box with the element symbol is the **atomic number** of the element which also represents the number of **protons** in each atom of the element.

The Periodic Chart is based on the properties of matter (see ***Properties of Matter*** in Unit 1), where a ***property*** may be defined as a quality or trait or characteristic. It is often possible to describe, identify, separate, and classify matter through its properties. For instance, it is possible to pick out a person from a small crowd of people based on a description (As long as the description is not too inaccurate, too vague, or too biased.). In a similar fashion, one can collect a number of properties to describe an element or compound. The properties of the element or compound, though, are true for any amount of the material anywhere, such that a South African diamond is indistinguishable from an Angolan diamond by its properties.

Periodic Properties

The periodic chart came about from the idea that we could arrange the elements, originally by atomic weight, in a scheme that would show similarity among groups. The original idea came from noticing how other elements combined with oxygen. Oxygen combines in some way with all the elements except the inert gases. Each atom of oxygen combines with two atoms of any element in Group 1, the elements in the row below lithium. Each atom of oxygen combines one-to-one with any element in Group 2, the elements in the row below beryllium. From here as we investigate the groups from left to right across the Periodic Chart, the story is not quite so clear, but the pattern is there. Group 3 is the group below boron. All of these elements combine with oxygen at the ratio of one-and-a- half to one oxygen. Group 4, beginning with carbon, combines two to one with oxygen. The group of transition elements (numbers 21-30 and 39-48 and 71-80 and 103 up) have never been adequately placed into the original scheme relating to oxygen. The transition elements vary in the ways they can attach to oxygen, but in a manner that is not so readily apparent by the simple scheme. Gallium, element number thirty-one, is the crowning glory of the Periodic Chart as first proposed by Mendeleev. Dmitri Ivanovich Mendeleev first proposed the idea that the elements could be arranged in a periodic fashion. He left a space for gallium below aluminum, naming it eka- aluminum, and predicting the properties of gallium fairly closely. The element was found some years later just as Mendeleev had predicted. Mendeleev also accurately predicted the properties of other elements.

Most Periodic Charts have two rows of fourteen elements below the main body of the chart. These two rows, the Lanthanides and Actinides really should be in the chart from numbers 57 - 70 and from 89 - 102. To show this, there would have to be a gulf of fourteen element spaces between numbers 20 - 21 and numbers 38 - 39. This would make the chart almost twice as long as it is now. The Lanthanides belong to Period 6, and the Actinides belong to Period 7. In basic Chemistry courses you will rarely find much use for any of the Lanthanides or Actinides, with the possible exception of Element #92, Uranium. No element greater than #92 is found in nature. They are all man-made elements, if you would like to call them that. None of the elements greater than #83 have any isotope that is completely stable. This means that all the elements larger than bismuth are naturally radioactive. The Lanthanide elements are so rare that you are not likely to run across them in most beginning chemistry classes. Another oddity of the Periodic Chart is that hydrogen does not really belong to Group I -- or any other group. Despite being over seventy percent of the atoms in the known universe, hydrogen is a unique element.

Symbols of Elements

For every element, there is one and only one upper case letter. There may or may not be a lower case letter with it. When written in chemical equations, we represent the elements by the symbol alone with no charge attached. The seven exceptions to that are the seven elements that are in gaseous form as a diatomic molecule, that is, two atoms of the same element attached to each other. The list of these elements is best memorized. They are: hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine. The chemical symbols for these diatomic gases are: H_2, N_2, O_2, F_2, C_{l2}, Br_2, and I_2. Under some conditions oxygen makes a triatomic molecule, ozone, O_3. Ozone is not stable, so the oxygen atoms rearrange themselves into the more stable diatomic form.

In determining the symbols of elements in the periodic chart:

- Use 1 or 2 letter(s) abbreviations.
- Capitalize the first letter only

Examples:

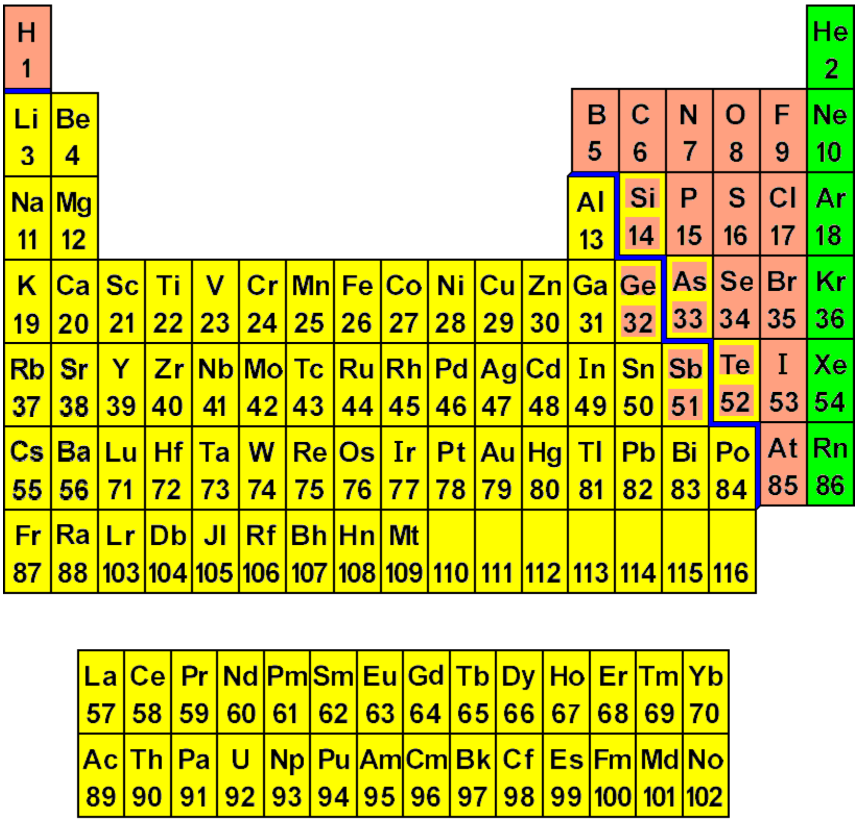
C carbon
N nitrogen
F fluorine
Co Cobalt
O Oxygen
Ca calcium
Br bromine
Mg magnesium

The symbols were derived from Latin names as shown in table below:

Element	**Symbol**	**Latin name**
Copper	Cu	cuprum
Gold	Au	aurum
Lead	Pb	plumbum
Mercury	Hg	hydrargyrum
Potassium	K	kalium
Silver	Ag	argentum
Sodium	Na	natrium
Tin	Sn	stannum

Note: For better learning, it is recommended that a few short lists be well learned for immediate recognition. These include: the diatomic gases (hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine), the Group one elements (lithium, sodium, potassium, rubidium, cesium, and francium), the Group two elements (beryllium, magnesium, calcium, strontium, barium, and radium), Group seven elements, the halogens, (fluorine, chlorine, bromine, iodine, and astatine), and the noble gases (helium, neon, argon, krypton, xenon, and radon). If nothing else, learning these as a litany will help you distinguish between radium, a Group 1 element, and radon, an inert gas.

Periodic Chart Of The Elements



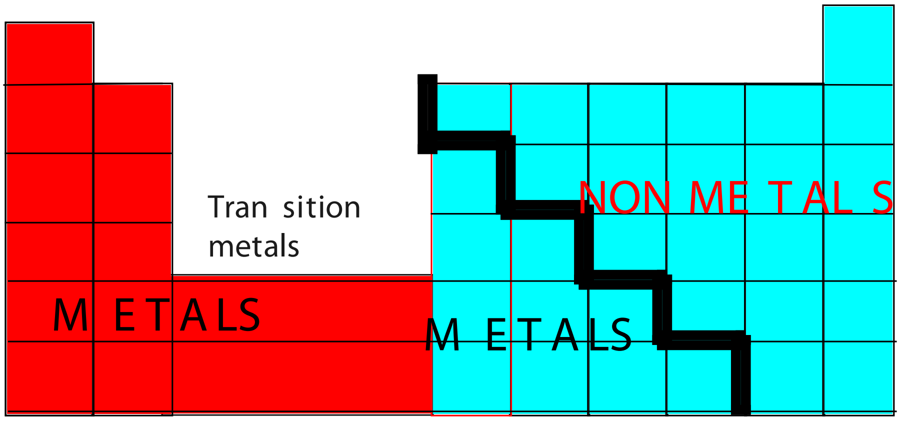
H 1																	He 2
Li 3	Be 4											B 5	C 6	N 7	O 8	F 9	Ne 10
Na 11	Mg 12											Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
K 19	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Br 35	Kr 36
Rb 37	Sr 38	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48	In 49	Sn 50	Sb 51	Te 52	I 53	Xe 54
Cs 55	Ba 56	Lu 71	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80	Tl 81	Pb 82	Bi 83	Po 84	At 85	Rn 86
Fr 87	Ra 88	Lr 103	Db 104	Jl 105	Rf 106	Bh 107	Hn 108	Mt 109	110	111	112	113	114	115	116		

La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70
Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102

Consider a staircase-shaped line on the Periodic Chart starting between boron and aluminum turns to be between aluminum and silicon then down between silicon and germanium, between germanium and arsenic, between arsenic and antimony, between antimony and tellurium, between tellurium and polonium, and between polonium and astatine. This is the line between metal and non-metal elements (see figure below). Metal elements are to the left and down from the line and non-metal elements are to the right and up from the line. Well, that is not exactly true. There is a line of non-metal elements, Group 8, or Group 18, or Group 0, whichever way you count them, the noble or inert gases that are really an entire Group and category to themselves. Hydrogen is a unique element, the only member of its own Group and category.

Figure below, gives the grouping of the various elements according to their properties:

Metals and Nonmetals of the Periodic Table



Groups or Families of the Periodic Chart

This section is not intended to be an exhaustive study of the groups of the Periodic Chart, but a quick overview of the groups as a way to see the organization of the chart.

Note: Many texts and charts will label the groups with different names and numbers. In here, an attempt is made to give some standard numbers and identify the elements in those groups so there is no question about which ones are being described. It is a good idea to refer to the Periodic Chart availed here as you go through this section.

Group I (1) elements, lithium, sodium, potassium, rubidium, cesium, and francium, are also called the alkali metal elements. They are all very soft metals that are not found free in nature because they react with water. In the element form they must be stored under kerosene to keep them from reacting with the humidity in the air. They all have a valence of plus one because they have one and only one electron in the outside shell. All of the alkali metals show a distinctive color when their compounds are put into a flame. The lightest of them are the least reactive. Activity increases as the element is further down the Periodic Chart. Lithium reacts leisurely with water. Cesium reacts very violently. Very few of the salts of Group 1 elements are not soluble in water. The lightest of the alkali metals are very common in the earth's crust. Francium is both rare and radioactive.

Group II (2) elements, beryllium, magnesium, calcium, strontium, barium, and radium, all have two electrons in the outside ring, and so have a valence of two. Also called the alkaline earth metals, Group 2 elements in the free form are slightly soft metals. Magnesium and calcium are common in the earth's crust.

Group 3 elements, boron, aluminum, gallium, indium, and thallium, are a mixed group. Boron has mostly non- metal properties. Boron will bond covalently by preference. The rest of the group are metals. Aluminum is the only one common in the earth's crust. Group 3 elements have three electrons in the outer shell, but the larger three elements have valences of both one and three.

Group 4 elements, carbon, silicon, germanium, tin, and lead, are not a coherent group either. Carbon and silicon bond almost exclusively with four covalent bonds. They both are common in the earth's crust. Germanium is a rare semi-metal. Tin and lead are definitely metals, even though they have four electrons in the outside shell. Tin and lead have some differences in their properties from metal elements that suggest the short distance from the line between metals and non-metals (semi-metal weirdness). They both have more than one valence and are both somewhat common in the earth's crust.

Group 5 is also split between metals and non-metals. Nitrogen and phosphorus are very definitely non-metals. Both are common in the earth's crust. In the rare instances that nitrogen and phosphorus form ions, they form triple negative ions. Nitride (N^{-3}) and phosphide (P^{-3}) ions are unstable in water, and so are not found in nature. All of the Group 5 elements have five electrons in the outer shell. For the smaller elements it is easier to complete the shell to become stable, so they are non-metals. The larger elements in the group, antimony and bismuth, tend to be metals because it is easier for them to donate the five electrons than to attract three more. Arsenic, antimony and bismuth have valences of +3 or +5. Arsenic is very much a semi-metal, but all three of them show some semi-metal weirdness, such as brittleness as a free element.

Group VII (6 or 16) elements, oxygen, sulfur, selenium, and tellurium, have six electrons in the outside shell. We are not concerned with polonium as a Group 6 element. It is too rare, too radioactive, and too dangerous for us to even consider in a basic course. Tellurium is the only element in Group 6 that is a semi-metal. There are positive and negative ions of Tellurium. Oxygen, sulfur, and selenium are true non-metals. They have a valence of negative two as an ion, but they also bond covalently. Oxygen gas makes covalent double-bonded diatomic molecules. Oxygen and sulfur are common elements. Selenium has a property that may be from semi-metal weirdness; it conducts electricity much better when light is shining on it. Selenium is used in photocells for this property.

On some charts you will see hydrogen above fluorine in Group VII (7 or 17). Hydrogen does not belong there any more than it belongs above Group 1. Fluorine, chlorine, bromine, and iodine make up Group 7, the halogens. We can forget about astatine. It is too rare and radioactive to warrant any consideration here. Halogens have a valence of negative one when they make ions because they have seven electrons in the outer shell. They are all diatomic gases as free elements near room temperature. They are choking poisonous gases. Fluorine and chlorine are yellow-green, bromine is reddish, and iodine is purple as a gas. All can be found attached to organic molecules. Chlorine is common in the earth's crust.

Fluorine is the most active of them, and the activity decreases as the size of the halogen increases.

The inert gases or noble gases all have a complete outside shell of electrons. Helium is the only one that has only an "*s*" subshell filled, having only two electrons in the outer and only shell. All the others, neon, argon, krypton, xenon, and radon, have eight electrons in the outer shell. Since the electron configuration is most stable in this shape, the inert gases do not form natural compounds with other elements. The group is variously numbered as Group VII, 8, 8A, 0, or 18. 'Group zero' seems to fit them nicely since it is easy to think of them as having a zero valence, that is no likely charge.

The Transition Elements make up a group between what has been labeled as Group 2 and Group 3. Transition elements are all metals. Very few of the transition elements have any non-metal properties. Within the transition elements many charts subdivide the elements into groups, but other than three horizontal groups, it is difficult to make meaningful distinctions among them. The horizontal groups are: iron, cobalt, and nickel; ruthenium, rhodium, and palladium; and osmium, iridium, and platinum.

Lanthanides, elements 57 through 70, are also called the rare earth elements. They are all metal elements very similar to each other, but may be divided into a cerium and a yttrium group. They are often found in the same ores with other elements of the group. None are found in any great quantity in the earth's crust. Of the Actinides, elements 89 through 102, only the first three are naturally occurring, the rest being manufactured elements. Of the three naturally occurring ones, only uranium is likely to be referred to in any way in a basic chemistry course. Elements 103 through 109 have been manufactured, and they have been named by the IUPAC (International Union of Pure and Applied Chemistry), but they are not of much importance to this course.

Formative evaluation :

Exercise 1: Without looking at the periodic table, identify the group numbers, and periodic number of an atom that has the following configuration:

(a) [Ne] $3s^1$ (b) [He] $2s^2$ (c) [Kr] $5s^2$ $4d^{10}$ $5p^5$

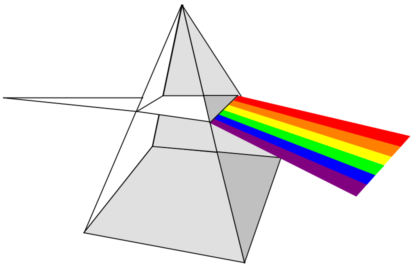
Exercise 2: For each of the elements below write the full and condensed electron configuration:

a) Potassium (b) Carbon (c) Calcium.

Exercise 3: How do electron shells, subshells,and orbitals differ?

Exercise 4: Determine the number of protons and neutrons in the nucllei of each the following:

(a) N, atomic number 7
(b) Al, atomic number 13
(c) O, atomic number 8



Learning activity # 3

Title of Learning Activity: Bonding, Molecules and Periodicity

Specific Learning Objectives

- Distinguish between ionic and covalent bonding.
- Describe and explain the criteria for determining types of intermolecular bonding.
- Use IUPAC nomenclature to name and write formula for binary inorganic compounds, cations, and anions.
- Explain and predict chemical bonding, structure and geometrics of simple molecules and ions.
- Describe and apply Lewis theory in making Lewis structures of elements and binary compounds.

Summary of the learning activity

The basic composition of different substances is atoms. Yet how different are the properties of the substances that we meet on a day to day basis. These properties are related to the kinds of particles of which the substances are composed. For instance, the water in a bowl of soup consists of molecules while the common salt in a shaker is composed of positive and negative ions. Why elements interact to form these different substances is the subject of this unit. This unit looks at the concept behind the arrangement of atoms of various elements to form new entities that themselves contain the individual atoms, linked together in some fashion. The nature of the interactions between the various atoms is considered in depth. Factors influencing the formation of ionic versus molecular compounds are also covered. Concepts behind the formation of different types of bonding (e.g., ionic versus covalent) are covered in this unit. Also, the international system of nomenclature adopted in naming and writing the formulas of simple binary inorganic and organic compounds, cations and anions are considered. A section of the module deals with explaining and predicting the nature of a chemical bond, structure and geometry of simple molecules and ions. A description and application of Lewis theory in making Lewis structures of elements and binary compounds is also covered. In here, the procedure followed in drawing Lewis structures is also covered.

Key Concepts

A bond: is an attachment among atoms.

Anion: a negative ion.

Cation: a positive ion, whether it is a single atom or a group of atoms all with the same charge.

Compound: a group of atoms with an exact number and type of atoms in it arranged in a specific way.

Molecule: the smallest mass of an element or compound capable of existing alone and possessing the properties of that element or compound.

Covalent bond: a bond that results from the sharing of a pair of electrons between atoms.

Covalent compounds: groups of atoms joined by covalent bonds.

Ionic bond: the attraction between a positive ion and a negative ion constitutes an ionic bond.

Ionic compound: a group of atoms attached by an ionic bond that is a major unifying portion of the compound.

Valence: the likely charge that an atom will take on as an ion.

List of relevant readings

Atomic Structure.htm
Chemical Nomenclature.htm
Review of Elements, Compounds, and Mixtures.htm
Nomen.htm
Periodicity.htm
Notes Molecules & Compounds.htm
Naming of Inorganic Compounds.htm
Naming Compounds, Ions, and Acids.htm

List of relevant useful links

http://www.chemtutor.com/compoun.htm

http://www.chemtutor.com/peric.htm

Detailed description of the activity

Ionic and Covalent Bonding

A ***bond*** is an attachment among atoms. Atoms may be held together for any of several reasons, but all bonds have to do with the electrons, particularly the outside electrons of atoms. The different type of bonds may be categorized as follows:

- There are bonds that occur due to sharing electrons.
- There are bonds that occur due to a full electrical charge difference attraction.
- There are bonds that come about from partial charges or the position or shape of electrons about an atom.

The bottomline, however, is that **all** bonds have to do with electrons.

Since chemistry is the study of elements, compounds, and how they change, it might be said quite loosely that chemistry is the study of electrons. In chemical reactions, the elements involved do not change from one element to another, but are only rearranged in their attachments.

A ***compound*** is a group of atoms with an exact number and type of atoms in it arranged in a specific way. Every bit of that material is exactly the same. Exactly the same elements in exactly the same proportions are in every bit of the compound. An example of a compound is water. In this, one oxygen atom and two hydrogen atoms make up water. Each hydrogen atom is attached to an oxygen atom by a bond. Any other arrangement is not water and if any other elements are attached, it ceases to be water. H_2O is the formula for the compound water. This formula indicates that there are two hydrogen atoms and one oxygen atom in the compound. H_2S is hydrogen sulfide and it does not have the same types of atoms as water. Hence, it is a different compound. H_2O_2 is the formula for hydrogen peroxide. It might have the right elements in it to be water, but it does not have them in the right proportion and so It is still not water. The word formula is also used to mean the smallest bit of any compound. A **molecule** is a single formula of a compound joined by **covalent bonds**.

The Law of Constant Proportions states that a given compound always contains the same proportion by weight of the same elements.

Ionic Bonds

Some atoms, such as metals tend to lose electrons to make the outside ring or rings of electrons more stable and other atoms tend to gain electrons to complete the outside ring. An ***ion*** is a charged particle. Electrons are negative. The negative charge of the electrons can be offset by the positive charge of the protons, but the number of protons does not change in a chemical reaction. When an atom loses electrons it becomes a positive ion because the number of protons exceeds the number of electrons. Non-metal ions and most of the polyatomic ions have a negative charge. The non-metal ions tend to gain electrons to fill out the outer shell. When the number of electrons exceeds the number of protons, the ion is negative. The attraction between a positive ion and a negative ion constitutes an **ionic bond**. Any positive ion will bond with any negative ion (opposites attract).

An **ionic compound** is a group of atoms attached by an ionic bond that is a major unifying portion of the compound. A positive ion, whether it is a single atom or a group of atoms all with the same charge, is called a ***cation***. A negative ion is called an ***anion***. The name of an ionic compound is the name of the positive ion (cation) first and the negative (anion) ion second.

The ***valence*** of an atom is the likely charge that it will take on as an ion. The names of the ions of metal elements with only one valence, such as the Group 1 or Group 2 elements, is the same as the name of the element. The names of the ions of nonmetal elements (anions) develop an -ide on the end of the name of the element. For instance, fluorine ion is fluoride, oxygen ion is oxide, and iodine ion is iodide. There are a number of elements, usually transition elements that having more than one valence, have a name for each ion. For instance, ferric ion is an iron ion with a positive three charge whereas ferrous ion is an iron ion with a charge of plus two.

There are also a number of common groups of atoms that have a charge for the whole group. Such a group is called a polyatomic ion or radical. It is suggested that the learner learns by rote, the list of polyatomic ions with their corresponding names, formulas and charges.

Writing Ionic Compound Formulas

Knowing the ions is the best way to identify ionic compounds and to predict how materials would join. Those who do not know of the ammonium ion and the nitrate ion would have difficulty seeing that NH_4NO_3 is ammonium nitrate. The writer of this module therefore, **very highly** recommends that the learner knows all the ions listed in the relevant reference provided below, complete with the valence or charge.

Note: The reader is hereby referred to subsections in the file "CHEMTUTOR COMPOUNDS.htm" titled "SOME ATOMS WITH MULTIPLE VALENCES"; "SOME ATOMS WITH ONLY ONE COMMON VALENCE"; "RADICALS OR POLYATOMIC IONS"; and "ACIDS OF SOME COMMON POLYATOMIC IONS".

In the reference sections indicated above, the learner will notice that the radicals and compounds have a small number after and below an element if there is more than one of that type of that atom. For instance, ammonium ions have one nitrogen atom and four hydrogen atoms in them. Sulfuric acid has two hydrogens, one sulfur, and four oxygens.

Now, consider what happens in an ionic bond using **electron configuration**, the **octet rule**, and some creative visualization. A sodium atom has eleven electrons around it. The first **shell** has two electrons in an ***s subshell***. The second shell is also full with eight electrons in an ***s*** and a ***p*** subshell. The outer shell has one lonely electron, as do the other elements in **Group 1**. This outside electron can be detached from the sodium atom, leaving a sodium ion with a single positive charge and an electron.

On the other hand, a chlorine atom has seventeen electrons. Two are in the first shell, eight are in the second shell, and seven are in the outside shell. The outside shell is lacking one electron to make a full shell, as are all the elements of Group 7. When the chlorine atom collects another electron, the atom becomes a negative ion. The positive sodium ion missing an electron is attracted to the negative chloride ion with an extra electron. The symbol for a single unattached electron is e^-.

$\frac{1}{2}Cl_2 + Na \longrightarrow Cl + e^- + Na^+ \longrightarrow Cl^- + Na^+ \longrightarrow Na^+Cl^- \longrightarrow NaCl$

Any compound should have a net zero charge. The single positive charge of the sodium ion cancels the single negative charge of the chloride ion. The same idea would be for an ionic compound made of ions of plus and minus two or plus and minus three, such as magnesium sulfate or aluminum phosphate.

$Mg^{2+} + (SO_4)^{2-} \longrightarrow Mg^{2+}(SO_4)^{2-} \longrightarrow Mg(SO_4)$ or $MgSO_4$

$Al^{3+} + (PO_4)^{3-} \longrightarrow Al^{3+}(PO_4)^{3-} \longrightarrow Al(PO_4)$ or $AlPO_4$

The question however, is, what happens if the amount of charge does not match? For instance, Aluminum bromide has a cation that is triple positive and an anion that is single negative. The compound must be written with one aluminum and three bromide ions. $AlBr_3$. Calcium phosphate has a double positive cation and a triple negative anion. If you like to think of it this way, the number of the charges must be switched to the other ion. $Ca_3(PO_4)_2$. Note that there must be two phosphates in each calcium phosphate, so the parentheses must be included in the formula to indicate that. Each calcium phosphate formula (Ionic compounds do not make molecules.) has three calcium atoms, two phosphate atoms, and eight oxygen atoms.

There are however, a small number of ionic compounds that do not fit into the system for one reason or other. A good example of this is magnetite, an ore of iron, Fe_3O_4. The calculated charge on each iron atom would be +8/3, not a likely actual charge. The deviance from the system in the case of magnetite could be accounted for by a mixture of the common ferric and ferrous ions.

Binary Covalent Compounds

The word **binary** means that there are two types of atoms in a **compound**. **Covalent compounds** are groups of atoms joined by covalent bonds. Binary covalent compounds are some of the very smallest compounds attached by covalent bonds. A **covalent bond** is the result of the sharing of a pair of electrons between two atoms. The chlorine molecule is a good example of the bond, even if it has only one type of atom. Chlorine gas, Cl_2, has two chlorine atoms, each of which has seven electrons in the outside ring. Each atom contributes an electron to an electron pair that make the covalent bond. Each atom shares the pair of electrons. In the case of chlorine gas, the two elements in the bond have exactly the same pull on the electron pair, so the electrons are exactly evenly shared. The covalent bond can be represented by a pair of dots between the atoms, Cl:Cl, or a line between them, Cl-Cl. Sharing the pair of electrons makes each chlorine atom feel as if it has a completed outer shell of eight electrons. The covalent bond is much harder to break than an ionic bond. The ionic bonds of soluble ionic compounds come apart in water, but covalent bonds do not usually come apart in water. Covalent bonds make real molecules, groups of atoms that are genuinely attached to each other.

Binary covalent compounds have two types of atom in them, usually non-metal atoms. Covalent bonds can come in double (sharing of two pairs of electrons) and triple (three pairs of electrons) bonds.

The table below shows common names and their corresponding formula of some commonly encountered binary covalent compounds.

Formula	**Common name**	**System name**
N_2O	nitrous oxide	dinitrogen monoxide
NO	nitric oxide	nitrogen monoxide
N_2O_3	nitrous anhydride	dinitrogen trioxide
NO_2	nitrogen dioxide	nitrogen dioxide
N_2O_4	nitrogen tetroxide	dinitrogen tetroxide
N_2O_5	nitric anhydride	dinitrogen pentoxide
NO_3	nitrogen trioxide	nitrogen trioxide

From the above table, we see that there are often more ways for any two elements to combine with each other by covalent bonds than by ionic bonds. Many of the frequently seen compounds already have names that have been in use for a long time. These names, called common names, may or may not have anything to do with the makeup of the material, but more of the common names of covalent compounds are used than of the ionic compounds.

Note: that the ***system names*** include numbers that indicate how many of each type of atom are in a covalent molecule: mono- (one); di- (two); tri- (three); tetra- (four); penta- (five); hexa- (six); hepta- (seven); etc.

The mono-; di-; tri-; tetra-, etc are referred to as the Fake Greek Prefixes (FGPs) and are used to indicate the numbers.

In saying or writing the name of a binary covalent the FGP of the first element is said, then the name of the first element is said, then the FGP of the second element is said, and the name of the second element is said, usually with the ending "-ide" on it. The only notable exception for the rule is if the first mentioned element only has one atom in the molecule, in which case the "mono-" prefix is omitted. CO is carbon monoxide. CO_2 is carbon dioxide. In both cases there is only one carbon in the molecule, and the "mono-" prefix is not mentioned. For oxygen, the last vowel of the FGP is omitted, as in the oxides of nitrogen in the above table.

Common names of binary covalent compounds that the learner is expected to know are: H_2O, Water; NH_3, ammonia; N_2H_4, hydrazine; CH_4, methane; and C_2H_2, acetylene.

Characteristics of Ionic and Covalent Compounds

Ionic Compounds	Covalent Compounds
Contain positive and negative ions (Na^+Cl^-)	Exist as neutral molecules ($C_6H_{12}O_2$)
Solids such as table salt ($NaCl_{(s)}$)	Solids, liquids,or gases ($C_6H_{12}O_{6(s)}$, $H_2O_{(l)}$, $CO_{2(g)}$)
High melting and boiling points	Lower melting and boiling points (i.e., often exist as a liquid or gas at room temperature)
Strong force of attraction between particles	Relatively weak force of attraction between molecules
Separate into charged particles in water to give a solution that conducts electricity $NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$	Remain as same molecule in water and will not conduct electricity $C_6H_{12}O_{6(s)} \xrightarrow{H_2O} C_6H_{12}O_{6(aq)}$

Exercise 1: For each of the following compounds, predict whether you would expect it to be ionic or covalent.

a) chromium(III) oxide, Cr_2O_3
b) carbon tetrachloride, CCl_4
c) methanol, CH_3OH
d) strontium fluoride, SrF_2

Exercise 2: Use the following data to propose a way of distinguishing between ionic and covalent compounds.

Compound	**Melting Point (^oC)**	**Boiling Point (^oC)**
Cr_2O_3	2266	4000
SrF_2	1470	2489
CCl_4	-22.9	76.6
CH_3OH	-97.8	64.7

More on Bonds, Shapes, and Other forces

A covalent bond is a shared pair of electrons. The bond between the two atoms of any diatomic gas, such as chlorine gas, Cl_2, is certainly equally shared. The two chlorine atoms have exactly the same pull on the pair of electrons, so the bond must be exactly equally shared. In cesium fluoride the cesium atom certainly donates an electron and the fluoride atom certainly craves an electron. Both the cesium ion and the fluoride ion can exist independently of each other. The bond between a cesium and a fluoride ion is clearly ionic.

The amount of pull on an atom has on a shared pair of electrons, called **electronegativity**, is what determines the type of bond between atoms. Considering the **Periodic Chart** without the inert gases, electronegativity is greatest in the upper right of the Periodic Chart and lowest at the bottom left. The bond in francium fluoride should be the most ionic. Some texts refer to a bond that is between covalent and ionic as a polar covalent bond. There is a range of bond between purely ionic and purely covalent that depends upon the electronegativity of the atoms around that bond. If there is a large difference in electronegativity, the bond has more ionic character. If the electronegativity of the atoms is more similar, the bond has more covalent character.

Lewis Structures

Lewis structures are an opportunity to better visualize the valence electrons of elements. In the **Lewis model**, an element symbol is inside the valence electrons of the ***s*** and ***p*** subshells of the outer ring.

Note: It is not very convenient to show the Lewis structures of the Transition Elements, the Lanthanides, or Actinides and are beyond the scope of this module.

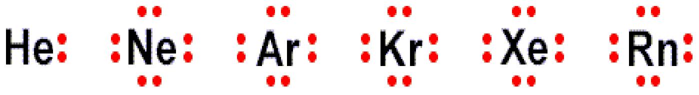
The inert gases are shown having the element symbol inside four groups of two electrons symbolized as dots. Two dots above the symbol, two below, two on the right, and two on the left. The inert gases have a full shell of valence electrons, so all eight valence electrons appear. Halogens have one of the dots missing. It does not matter on which side of the symbol the dot is missing. Group 1 elements and hydrogen are shown with a single electron in the outer shell. Group 2 elements are shown with two electrons in the outer shell, but those electrons are not on the same side. Group 3 elements have three dots representing electrons, but the electrons are spread around to one per position, as in Group 2 elements. Group 4 elements, carbon, silicon, etc. are shown as having four electrons around the symbol, each in a different position.

Group 5 elements, nitrogen, phosphorus, etc. have five electrons in the outer shell. In only one position are there two electrons. So Group 5 elements such as nitrogen can either accept three electrons to become a triple negative ion or join in a covalent bond with three other items. When all three of the unpaired electrons are involved with a covalent bond, there is yet another pair of electrons in the outside shell of Group 5 elements.

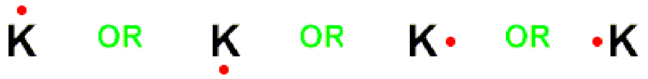
Group 6 elements, oxygen, sulfur, etc., have six electrons around the symbol, again without any concern to position except that there are two electrons in two positions and one electron alone in the other two positions. Group 7 elements have all of the eight outside electrons spaces filled except for one. The Lewis structure of a Group 7 element will have two dots in all four places around the element symbol except for one.

Lewis Structures of the Elements

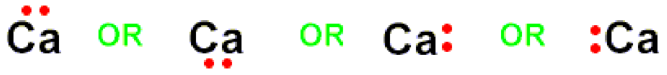
Examine the Sideways Periodic Chart With Electron Shell Numbers again. All of the Group I elements and hydrogen (the top row of the chart) have one and only one electron in the outside shell. That single electron is what gives these elements the distinctive character of the group. The Lewis structures are just an attempt to show these valence electrons in a graphic manner as they are used to combine with other elements. The element symbol is in the center and as many as four groups of two electrons are shown as dots above, below, to the right and left of the element symbol to show the valence electrons. All of the inert gases (noble gases) have all eight of the electrons around the element symbol, except for helium, which has only two electrons even with a full shell. Below is a demonstration of the noble gases written in Lewis structure. Notice the electrons are in red just to emphasize them.



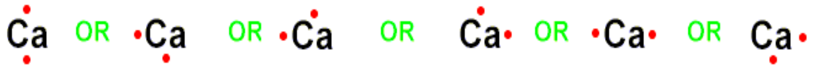
All the other elements have less than eight electrons in the outside shell. These electrons can be in the positions of the eight electrons of the noble gases, but there are some suggestions about where they belong. The Group I elements have only one electron in the outer shell, so it really does not matter where the electron dot is placed, over, under, right or left of the element symbol.



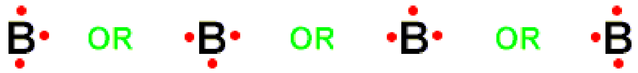
Group II elements have two electrons. Some authors will place the two electron dots together on any side of the element symbol because the electrons really are in an *s* subshell together.



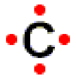
Some authors will show the electrons separated from each other in any of the two positions with only one electron in each position. The reasoning behind that is that the electrons really do try to move as far away from each other as possible.



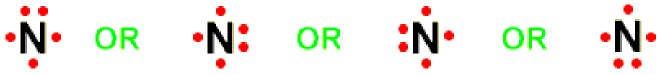
Boron and the elements below it on the periodic table all have three electrons in the outside shell. These electrons may be grouped as each electron alone in one of the positions around the element symbol or as a group of two (*s*) electrons in one position and one electron in another. Boron is usually shown with separate electrons because it bonds mostly covalently. Covalent bonds, we know from the shape of molecules, tend to blend the *s* and *p* subshells into *sp* orbitals with one *s* and one *p* orbital blended, *sp2* orbitals with one *s* and two *p* orbitals blended, or *sp3* orbitals, using the single *s* orbital with all three *p* orbitals. The *sp2* orbitals of boron tend to be flat trigonal shape, that is, the bonds are at 120 degrees from each other in a flat circle around the boron atom in the center. The Lewis structure of boron is any of the shapes below.



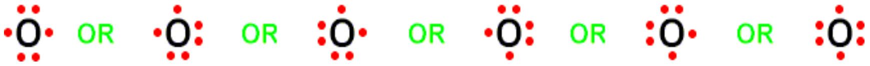
Carbon and the elements below it have four electrons in the outer shell. Carbon and silicon are usually shown in Lewis structures to have four separated electrons, again because these elements bond purely with covalent bonds. The *sp3* orbitals of carbon and silicon are tetrahedral in shape.



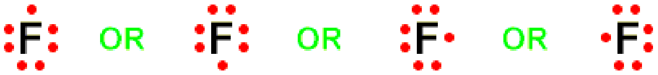
Nitrogen and the elements below it have five electrons in the valence shell, so they must be shown with one pair (anywhere) and three solitary electrons.



Oxygen and the elements below it have six valence electrons and so must have two pairs and two solitary electrons.

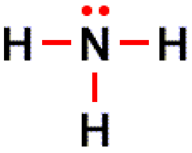


Elements in the halogen group, Group VII, all have seven electrons in the outer shell, so only there are three groups of two and a single electron in the last position.



The transition elements and the Lanthanide and Actinide series elements are not often used in the covalent bonds that the Lewis structures usually portray, but these metal elements can be portrayed in this manner using the number of electrons in the outer shell that corresponds with the valence of the element.

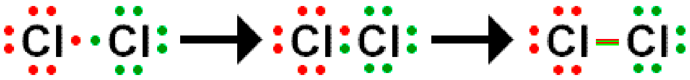
In using the Lewis structures to show covalent bonds, the pair of electrons that are in the bond are shown as a dashed line. For example, ammonia would be shown with the bonds from the nitrogen to the hydrogens and the unshared pair of electrons on the nitrogen.



The covalent bond dash is in red in the above sketch also. Notice that the electrons from all of the participants in this molecule are all accounted for.

In this section, the emphasis will be on the Lewis structure of small compounds and polyatomic ions.

Consider two atoms of the same type sharing a pair of electrons. Chlorine atoms have seven electrons each and would be a lot more stable with eight electrons in the outer shell. Single chlorine atoms just do not exist because they get together in pairs to share a pair of electrons. The shared pair of electrons make a bond between the atoms. In Lewis structures, the outside electrons are shown with dots and covalent bonds are shown by bars.



This covalent bond between chlorine is one of the most covalent bonds known. Why? A **covalent bond** is the sharing of a pair of electrons. The two atoms on ether side of the bond are exactly the same, so the amount of «pull» of each atom on the electrons is the same, and the electrons are shared equally.

Now, let us consider a molecule in which the atoms bonded are not the same, but the bonds are balanced such as methane, CH_4. If there were just a carbon and a single hydrogen, the bond between them would not be perfectly covalent. In the CH_4 molecule, the four hydrogen atoms exactly balance each other out. The Lewis structure of methane does not have any electrons left over. The carbon started with four electrons and each hydrogen started with two electrons. Only the bars representing the shared pairs of electrons remain. The carbon now shares four pairs of electrons, so this satisfies the carbon's need for eight electrons in the outside shell. Each hydrogen has a single shared pair in the outside shell, but the outside shell of the hydrogen only has two electrons, so the hydrogen has a full outer shell also.

Carbons and hydrogens are easy to write in Lewis structures, because each carbon must have four attachments to it and each hydrogen must have one and only one attachment to it. When the bonds around a carbon atom go to four different atoms, the shape of the bonds around that carbon is roughly tetrahedral, depending upon what the materials are around the carbon. Carbons are also able to have more than one bond between the same two.

Consider the series ethane (C_2H_6), ethene (C_2H_4),(common name is ethylene), and ethyne (C_2H_2), (common name is acetylene).

In writing the Lewis structure of compounds, the bars representing bonds are preferred to the dots representing individual electrons.

The double bars between the carbons in ethylene, C=C, represent a double bond between the two carbons, that is four shared electrons to make a stronger attachment between the two carbons. The triple bars between the carbons of acetylene represent a triple covalent bond between those two carbons, three pairs of shared electrons between those carbons. Every carbon has four bonds to it showing a pair of electrons to make eight electrons in the outer shell. Each hydrogen has one and only one bond to it for two electrons in the outer shell. All of the outer shells are filled.

While we are doing this, notice that the Lewis structure of a molecule will show the shape of the molecule. All of the bonds in ethane are roughly the tetrahedral angle, so all of the hydrogens are equivalent. This is true. The bonds in acetylene make it a linear molecule. The bonds in ethylene are somewhat trigonal around the carbons, and the carbons can not twist around that bond as they can around a single bond, so that the molecule has a flat shape and the hydrogens are not equivalent. This is also true. (You will see this in the study of organic chemistry. This type of difference between the positions of the hydrogens is called ***cis - trans*** isomerism.)

General guidelines for the drawing of Lewis structures for more complex molecules or polyatomic ions.

- Write all the atoms in the material.
- Usually pick the atom type with the most number of possible bonds to it to be the central atom or group of atoms. In most organic compounds, carbon provides the main "skeleton" of the molecule.
- Arrange the other materials around the inner core according the formula of the material.
- Arrange the electrons or bonds around each atom according to how many it should have.

Formative evaluation

Exercise 1: Name the following binary covalent compounds:

1. CO ________________________________
2. CO_2 ________________________________
3. S_2F_6 ________________________________
4. P_4O_{10} ________________________________
5. N_2O_4 ________________________________
6. NCl_3 ________________________________
7. PBr_5 ________________________________
8. SiS_2 ________________________________
9. N_2F_4 ________________________________
10. $SeBr_2$ ________________________________

Exercise 2: Name the following ionic compounds:

a) NaBr ________________________________
b) $CaSO_4$ ________________________________
c) K_2S ________________________________
d) $Ni(NO_3)_2$ ________________________________
e) Mg_3N_2________________________________
f) $Fe_2(CO_3)_3$ ________________________________
g) Cr_2O_3 ________________________________
h) $Ti(ClO_4)_4$ ________________________________
i) $AlCl_3$ ________________________________
j) PbC_2O_4 ________________________________

Exercise 3. Write chemical formula of the following and show subscript numbers where needed.

1. hydrochloric acid ____________________
2. sodium chloride __________________
3. sodium hexafluoride _______________
4. strontium nitrate __________________
5. calcium chloride ___________________
6. acetic acid _____________________
7. phosphoric acid ____________________
8. ammonia ________________________
9. lithium sulfate _____________________
10. potassium chromate ______________

Answers to naming binary compounds (Exercise 1)

1. carbon monoxide
2. carbon dioxide
3. disulfur hexafluoride
4. tetraphosphorus decoxide
5. dinitrogen tetroxide
6. nitrogen trichloride
7. phosphorus pentabromide
8. silicon disulfide
9. dinitrogen tetrafluoride
10. selenium dibromide

Answers to problems in naming ionic compounds (Exercise 2)

1. sodium bromide
2. calcium sulfate
3. potassium sulfide
4. nickel(II) nitrate
5. magnesium nitride
6. iron(III) carbonate
7. chromium(III) oxide
8. titanium(IV) perchlorate
9. aluminum chloride
10. lead(II) oxalate

Answers to Exercise 3

1. HCl
2. $NaCl$
3. NaF_6
4. $Sr(NO_3)_2$
5. $CaCl_2$
6. $HC_2H_3O_2$
7. H_3PO_4
8. NH_3
9. Li_2SO_4
10. K_2CrO_4

Learning activity # 4

Title of Learning Activity: Chemical Reactions and Stoichiometry

Specific Learning Objectives:

- Explain the mole concept, and convert between grams, moles, and atoms and molecules.
- Determine mass percent composition of a sample from experimental data.
- Determine empirical formula of a compound from its mass percent composition.
- Understand the concept of basic stoichiometry and related calculations.
- Write balanced chemical equations from descriptions of chemical changes.
- Interpret chemical equations in terms of reactants and products.
- Classify chemical reactions as say, precipitation, neutralization, combustion, decomposition, etc.

Summary of the learning activity

This unit begins by describing the mole and dwells on the methodologies involved in the interconversion between grams, moles, atoms and molecules. The theory behind determinination of percent composition of a sample by mass from experimental data is covered in great depth. The concepts behind determination of empirical formalas of compounds from their mass percent compositions are also covered. A good portion of the unit deals with the concept of basic stoichiometry and related calculations. The learner will be expected to master the concept of writing balanced equations from descriptions of various chemical changes. A classification of chemical reactions as say, precipitation, neutralization, combustion, decompositon, etc, is also covered covered in detail and the learner is encouraged to master these.

List of relevant readings

The Mole1.htm
Reactions.htm
Chemical Reactions.htm
ChemTeam Stoichiometry Molar Ratios.htm
ChemTeam Stoichiometry Mole-Mass Problems.htm
ChemTeam Stoichiometry Mole-Mole Problems.htm

List of relevant useful links

http://www.chem.vt.edu/RVGS/ACT/notes/molar_mass.html

Provides practice in calculating molar masses.

http://www.chem.vt.edu/RVGS/ACT/notes/The_mole.htm

Provides power point review notes on the “Mole”.

http://www.chemtutor.com/mols.htm

HTML site for material on mols, percent and stoichiometry

http://www.chemtutor.com/react.htm

HTML site for material on chemical reactions.

Detailed description of the activity

What is a Chemical Change?

A chemical change occurs whenever compounds are formed or decomposed. When a chemical change occurs, **reactants** disapper, while **products** appear. A chemical change occurs whenever there is rearrangement of atoms that makes or breaks chemical bonds.

Writing chemical equations

The steps involved in writing chemical equations entails the following:

1. Write a sentence describing the reaction that includes the names and states of all reactants and products, and any special conditions the reaction requires.

   Example: Ozone gas decomposes into oxygen gas when exposed to ultraviolet light.

2. Separate reactants and products with a process symbol:

   Ozone gas → Oxygen gas

   The following are the typical process symbols often used by chemists and their implications:

   → reaction goes to completion; sometimes used for one-step reactions.

   É reaction is **reversible**; at equilibrium, reactants and products are mixed

   = net reaction; reaction occurs in several steps.

3. Catalysts are usually written on top of the process symbol.
4. Replace names with formulas; put "**plus**" signs between formulas

   O_3 gas → O_2 gas

5. Include symbols for physical states.

   $O_3 \text{ (g)} \rightarrow O_2 \text{ (g)}$

   (g) gas; (l) liquid; (s) solid; and (aq) aqueous, i.e., dissolved in water.

6. Balance the equation by adding coefficients so atoms and charge are conserved. Each side of a balanced equation must have:
   - the same number of atoms of each type
   - the same total charge

   $2\ O_3 \text{ (g)} \rightarrow 3\ O_2 \text{ (g)}$

Types of chemical reactions

Some reactions can be classified in more than one way and some reactions fit into none of the sample categories given below:

Type	Definition	Examples
Synthesis	a compound forms from simpler reactants	$C_2H_4(g) + H_2O(g) \rightarrow C_2H_5OH(g)$ $H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l)$
Decomposition	a compound breaks down into simpler products	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $2 H_2O(l) \rightarrow 2 H_2(g) + O_2(g)$
displacement	$AB + C \rightarrow A + BC$	$Zn(s) + 2 HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ $Zn(s) + CuCl_2(aq) \rightarrow ZnCl_2(aq) + Cu(s)$
Neutralization	H^+ transfer from acid to base	$NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$
Precipitation	a solid product from aqueous reactants	$BaCl_2(aq) + Na_2SO_4(aq) = BaSO_4(s) + 2 NaCl(aq)$ $AgNO_3(aq) + NaCl(aq) = AgCl(s) + NaNO_3(aq)$
Redox	electron transfer between reactants	$2 C_2H_6(g) + 7 O_2(g) \rightarrow 4 CO_2(g) + 6 H_2O(g)$

Chemical Formulas and Equations

The Mole

Even tiny samples of chemicals contain huge numbers of atoms, ions or molecules. For convenience, some kind of reference for a collection of a large number of these objects would be very useful (e.g. a «dozen» is a reference to a collection of 12 objects). In chemistry we use a unit called a ***mole*** (abbreviated ***mol***).

A **mole** (also known as Avogadro's Number) is the number that is used in making calculations involving atoms and molecules. Given the relatively microscopic size of the particles, it would be rather difficult or utterly pointless to calculate amounts in terms of numbers of atoms. So, we use the mole.

By definition, ***a mole** is the amount of matter that contains as many objects as the number of atoms in exactly 12 grams of carbon-12.* In other words, **a mole** is equal to (all together now) **6.022 x 10^{23}** atoms or molecules, which is known as the **Avogadro's number**.

Molar Mass

The **atomic mass** given for an element on the periodic table measured out in grams is equal to ***one mole of atoms*** of that element. ***Note that the mass of one atom in atomic mass units (amu) is numerically equal to the mass of one mole of the same atoms in grams (g).***

A single carbon-12 (i.e., ^{12}C) atom has a mass of 12 *amu*. A single ^{24}Mg atom has a mass of 24 *amu*, or twice the mass of a ^{12}C atom. Thus, one mole of ^{24}Mg atoms should have twice the mass as one mole of ^{12}C atoms. Since one mole of ^{12}C atoms weighs 12 grams (by definition), one mole of ^{24}Mg atoms must weigh 24 grams.

The mass in grams of 1 mole (mol) of a substance is called its ***molar mass***. Thus, the **Molar Mass (MM)** of elements and compounds is the mass, in grams, equal to the atomic and formula masses of those elements and compounds. The unit of Molar Mass is **grams/mole**.

NOTE: The molar mass (in grams) of any substance is always numerically equal to its formula weight (in amu).

For instance,

One H_2O molecule weighs 18.0 amu; 1 mol of H_2O weighs 18.0 grams.

One NaCl ion pair weighs 58.5 amu; 1 mol of NaCl weighs 58.5 grams

Interconverting masses, moles, and numbers of particles

Keeping track of units in calculations is necessary when interconverting masses and moles. This is formally known as **dimensional analysis**.

Example: Suppose you are asked to provide 1.5 moles of calcium chloride. What does this entail?

Solution:

Chemical formula of calcium chloride = $CaCl_2$

Molecular mass of Ca = 40.078 amu

Molecular mass of Cl = 35.453 amu

Therefore, the ***formula weight*** of $CaCl_2$ = (40.078) + 2(35.453) = 110.984 *amu* (remember, this compound is ionic, so there is no «molecular» weight).

Therefore, one mole of $CaCl_2$ would have a mass of 110.984 grams. So, 1.5 moles of $CaCl_2$ would be:

(1.5 ~~mole~~)(110.984 grams/~~mole~~) = 166.476 grams

Example 2: How many atoms is equivalent to 2.8 grams of gold?

Solution:

Molecular formula of gold is: Au

Molecular weight of Au = 196.9665 *amu*

Therefore, 1 mole of gold weighs 196.9665 grams. So, in 2.8 grams of gold we would have:

(2.8 ~~gram~~)(1 mole/196.9665 ~~gram~~) = 0.0142 mole

From Avogadro's number, we know that there are approximately 6.02×10^{23} atoms/mole. Therefore, in 0.0142 moles we would have:

(0.0142 ~~mole~~)(6.02×10^{23} atoms/~~mole~~) = 8.56×10^{21} atoms

Example 3: If we have 2.00 mol of N_2 reacting with sufficient H_2, how many moles of NH_3 will be produced?

Solution:

1. The ratio from the problem will have N_2 and NH_3 in it.
2. How do you know which number goes on top or bottom in the ratios? Answer: it does not matter, except that you observe the next point ALL THE TIME.
3. When making the two ratios, be 100% certain that numbers are in the same relative positions. For example, if the value associated with NH_3 is in the numerator, then MAKE SURE it is in both numerators.

4. Use the coefficients of the two substances to make the ratio from the equation.
5. Why isn't H_2 involved in the problem? Answer: The word «sufficient» removes it from consideration.

Let's use this ratio to set up the proportion: $\frac{NH_3}{N_2}$

That means the ratio from the equation is: $\frac{2}{1}$

The ratio from the data in the problem will be: $\frac{x}{2.00}$

The proportion (setting the two ratios equal) is: $\frac{x}{2.00} = \frac{2}{1}$

Solving by cross-multiplying gives x = 4.00 mol of NH_3 produced.

Example 4: Suppose 6.00 mol of H_2 reacted with sufficient nitrogen. How many moles of ammonia would be produced?

Solution:

Let's use this ratio to set up the proportion: $\frac{NH_3}{H_2}$

That means the ratio from the equation is: $\frac{2}{3}$

The ratio from the data in the problem will be: $\frac{x}{6.00}$

The proportion (setting the two ratios equal) is: $\frac{x}{6.00} = \frac{2}{3}$

Solving by cross-multiplying and dividing gives x = 4.00 mol of NH_3 produced.

Empirical and Molecular Formulas

There exists more than one kind of chemical formula. From the above heading, one would be tempted to assume that only empirical and molecular formulas exist. **Empirical formulas** *give the lowest whole number ratio of the atoms in a compound*, whereas the **molecular formula** *gives the exact composition of one molecule*.

Illustration:

Empirical	**Molecular**
HO	H_2O_2
CH	C_2H_2
CH	C_6H_6
NO	NO
NO_2	N_2O_4

Calculating Empirical formulas using experimental data

Example: Suppose a given compound is 69.58% Ba, 6.090% C and 24.32% O, calculate the empirical formula of this compound. Assume that you have 100.00 g of the compound.

- 69.58% Ba = 69.58 g Ba; and 6.090%C = 6.09 g C; and 24.32% O = 24.32 g O.
- Now convert the mass of each element to moles of that element:

  moles of a substance = mass of substance (g) / molecular mass of substance

  Ba = (69.58 g Ba) (1 mol Ba /137.3 g Ba) = 0.5068 mol Ba

  C = (6.090 g C) (1 mol C/12.01 g C) = 0.5071 mol C

  O = (24.32 g O) (1 mol O/16.00 g O) = 1.520 mol O
- Divide through each value by the smallest number of moles to get a 1: 1.001: 2.999 ratio, which rounds off nicely to give the formula $BaCO_3$.

Calculating Molecular formulas using empirical formulas

Example: Given that the empirical formula of a compound is CH and the molar mass is 104 g/mol, calculate the molecular formula.

Mass of C = 12.0 g/mol

Mass of H = 1.01 g/mol

Empirical formula mass = 13.0 g/mol

- To find the number of CH units in the compound:

Number of CH units = (104 g/mol) (1 mol/13.0 g) = 8.00

- Molecular formula = 8(CH) or C_8H_8

What is Stoichiometry?

One can compare the amounts of any materials in the same chemical equation using the formula weights and the coefficients of the materials in a chemical equation. Consider the equation for the Haber reaction, which involves the combination of nitrogen gas and hydrogen gas to make ammonia.

$N_2 + 3\ H_2 \longrightarrow 2\ NH_3$

The formula for nitrogen is N_2 and the formula for hydrogen is H_2. They are both diatomic gases. The formula for ammonia is NH3. The balanced equation requires one (1) nitrogen molecule and three (3) hydrogen molecules to make two (2) ammonia molecules, meaning that one nitrogen molecule reacts with three hydrogen molecules to make two ammonia molecules or one MOL of nitrogen and three MOLS of hydrogen make two MOLS of ammonia.

Now, the real way of measuring amounts is by weight (actually, mass), so that 28 grams (14 g/mol times two atoms of nitrogen per molecule) of nitrogen and 6 grams of hydrogen (1 g/mol times two atoms of hydrogen per molecule times three mols) make 34 grams of ammonia. Notice that no mass is lost or gained, since the formula weight for ammonia is 17 (one nitrogen at 14 and three hydrogens at one g/mol) and there are two mols of ammonia made. Once you have the mass proportions, any **mass-mass stoichiometry** can be done by good old proportionation.

For instance, What is the mol ratio of hydrogen to ammonia in the Haber equation? 2 mols of hydrogen to 3 mols of ammonia. This should be easy. In the standard stoichiometry calculations you should know, ALL ROADS LEAD TO MOLS. You can change any amount of any measurement of any material in the same equation with any other material in any measurement in the same equation. That is powerful.

(1) Start with what you know (GIVEN), expressing it as a fraction.
(2) Use definitions or other information to change what you know to mols of that material.
(3) Use the **mol ratio** to exchange **mols** of the material given to the mols of material you want to find.
(4) Change the mols of material you are finding to whatever other measurement you need.

Example 1: How many grams of ammonia can you make with 25 grams of hydrogen? (Practice your mol math rather than doing this by proportion. Check it by proportion in problems that permit it.)

Solution:

You are given the mass of 25 grams of hydrogen. Start there.

$$\underset{\text{Given}}{\left(\frac{25\text{g } H_2}{1}\right)}\underset{\text{Fw}}{\left(\frac{\text{mol}}{2.0\text{g}}\right)}\underset{\text{Mol Ratio}}{\left(\frac{2 \text{ mols } NH_3}{3 \text{ mols } H_2}\right)}\underset{\text{Fw}}{\left(\frac{17\text{g}}{\text{mol}}\right)} = 142 \text{ g } NH_3$$

25 g H_2/1 change to mols of hydrogen by the formula weight of hydrogen, i.e., 1 mol of H_2 = 2.0 g. (The 2.0 g goes in the denominator to cancel with the gram units in the material given.) Now change mols of hydrogen to mols of ammonia by the mol ratio, i.e., 3 mols of hydrogen = 2 mols of ammonia. (The mols of hydrogen go in the denominator to cancel with the mols of hydrogen. You are now in units of mols of ammonia.) Convert the mols of ammonia to grams of ammonia by the formula weight of ammonia, i.e., 1 mol of ammonia = 17 g. (Now the mols go in the denominator to cancel with the mols of ammonia.) Cancel the units as you go.

Example 2: Calculate the mass of AgCl that can be prepared from 200. g of $AlCl_3$ and sufficient $AgNO_3$, using this equation: 3 $AgNO_3$ + $AlCl_3$ --> 3 AgCl + $Al(NO_3)_3$

Solution:

$\frac{AlCl_3}{AgCl}$	$\frac{200 \text{ g}}{133.34 \text{ g/mol}} = 1.50 \text{ mol}$
$\frac{1.50 \text{ mol}}{x} = \frac{1}{3}$	4.50 mol x 143.32 g/mol = 645 g

Example 3: Consider the reaction: $2\ H_2 + O_2$ ---> $2\ H_2O$.

(a) What is the molar ratio between H_2 and O_2?
(b) What is the molar ratio between O_2 and H_2O?

Solution: The molar ratio between H_2 and O_2 is 2:1. Ratio in fractional form is: $\frac{2}{1}$. It is recommended that you explicitly write the 1 in the denominator of the ratio.

The molar ratio between O_2 and H_2 is 1:2. Ratio in fractional form is: $\frac{1}{2}$.

Exercise 1: Write the molar ratios for N2 to H2 and NH3 to H2 in the equation:

$N_2 + 3\ H_2$ ---> $2\ NH_3$

Exercise 2: Write the molar ratios for O_2 to SO_3 and O_2 to SO_2 from the equation:

$2\ SO_2 + O_2$ ---> $2\ SO_3$

The word **stoichiometry** derives from two Greek words: *stoicheion* (meaning «element») and *metron* (meaning «measure»). **Stoichiometry** deals with calculations about the masses (sometimes volumes) of reactants and products involved in a chemical reaction. It is a very mathematical part of chemistry, so be prepared for lots of calculator use.

According to Jeremias Benjaim Richter, the first to lay down the principles of stoichiometry, *stoichiometry is the science of measuring the quantitative proportions or mass ratios in which chemical elements stand to one another.*

The most common stoichiometric problem will present you with a certain amount of a reactant and then ask how much of a product can be formed.

Consider a hypothetical chemical equation represented by:

2 A + 2B ---> 3C

Suppose you are informed that you have 20.0 grams of reactant A and sufficient amount of reactant B, how many grams of C can be produced?

To solve such a problem, you will need to use molar ratios, molar masses, balancing and interpreting equations, and conversions between grams and moles.

Note: If you struggled with those in the preceding sections, then welcome to the club. You need to go back and review them if you need to, because if you can't do that stuff, you will find it difficult to do stoichiometry.

This type of a problem is often called “mass-mass”.

The Steps Involved in Solving Mass-Mass Stoichiometry Problems

1. Make sure the chemical equation is correctly balanced.
2. Using the molar mass of the given substance, convert the mass given in the problem to moles.
3. Construct a molar proportion (two molar ratios set equal to each other) following the guidelines set out in other files. Use it to convert to moles of the unknown.
4. Using the molar mass of the unkown substance, convert the moles just calculated to mass.

Formative evaluation

Exercise 1. Write the formula for each material correctly and then balance the equation. There are some reactions that require completion. For each reaction, indicate what type of reaction it is.

a) sulfur trioxide and water combine to make sulfuric acid.
b) lead II nitrate and sodium iodide react to make lead iodide and sodium nitrate.
c) calcium fluoride and sulfuric acid make calcium sulfate and hydrogen fluoride (Hydrofluoric acid)
d) calcium carbonate will come apart when you heat it to leave calcium oxide and carbon dioxide.
e) ammonia gas when it is pressed into water will make ammonium hydroxide.
f) sodium hydroxide neutralizes carbonic acid.
g) zinc sulfide and oxygen become zinc oxide and sulfur.
h) lithium oxide and water make lithium hydroxide.
i) aluminum hydroxide and sulfuric acid neutralize to make water and aluminum sulfate.
j) sulfur burns in oxygen to make sulfur dioxide.
k) barium hydroxide and sulfuric acid make water and barium sulfate.
l) aluminum sulfate and calcium hydroxide become aluminum hydroxide and calcium sulfate.
m) copper metal and silver nitrate react to form silver metal and copper II nitrate.
n) sodium metal and chlorine react to make sodium chloride.
o) calcium phosphate and sulfuric acid make calcium sulfate and phosphoric acid.
p) phosphoric acid plus sodium hydroxide.
q) propane burns (with oxygen).

r) zinc and copper II sulfate yield zinc sulfate and copper metal.
s) sulfuric acid reacts with zinc.
t) acetic acid ionizes.

Solutions to above Questions

(a)	$SO_3 + H_2O \longrightarrow H_2SO_4$	(Synthesis)
(b)	$Pb(NO_3)_2 + 2NaI \longrightarrow PbI_2 \downarrow + 2NaNO_3$	(Double Replacement) (lead II iodide precipitates)
(c)	$CaF_2 + H_2SO_4 \longrightarrow CaSO_4 \downarrow + 2\ HF$	(Double Replacement) (calcium sulfate precipitates)
(d)	$CaCO_3 \longrightarrow CaO + CO_2$	(Decomposition)
(e)	$NH_3 + H_2O \longrightarrow NH_4OH$	(Synthesis)
(f)	$2\ NaOH + H_2CO_3 \longrightarrow Na_2CO_3 + 2\ H_2O$	(Double Replacement Or Acid-base Neutralization)
(g)	$2\ ZnS + O_2 \longrightarrow 2\ ZnO + 2\ S$	(Anionic Single Replacement)
(h)	$Li_2O + H_2O \longrightarrow 2\ LiOH$	(Synthesis)
(i)	$2\ Al(OH)_3 + 3\ H_2SO_4 \longrightarrow 6\ H_2O + Al_2(SO_4)_3$	(Double Replacement or Acid-base Neutralization)
(j)	$S + O_2 \longrightarrow SO_2$	(Synthesis)
(k)	$Ba(OH)_2 + H_2SO_4 \longrightarrow 2\ H_2O + BaSO_4$	(Double Replacement or Acid-base Neutralization)
(l)	$Al_2(SO_4)_3 + 3\ Ca(OH)_2 \longrightarrow 2\ Al(OH)_3 \downarrow + 3\ CaSO_4 \downarrow$	(Double Replacement - Both calcium sulfate and aluminum hydroxide are precipitates.)
(m)	$Cu + 2AgNO_3 \longrightarrow 2Ag + Cu(NO_3)_2$	(Cationic Single Replacement)
(n)	$2Na + Cl_2 \longrightarrow 2\ NaCl$	(Synthesis)
(o)	$Ca_3(PO_4)_2 + 3\ H_2SO_4 \longrightarrow 3\ CaSO_4 + 2\ H_3PO_4$	(Double Replacement)
(p)	$H_3(PO_4) + 3\ NaOH \longrightarrow Na_3PO_4 + 3\ H_2O$	(Double Replacement (Neutralization))
(q)	$C_3H_8 + 5\ O_2 \longrightarrow 4\ H_2O + 3\ CO_2$	(Burning Of A Hydrocarbon)
(r)	$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$	(Cationic Single Replacement)
(s)	$H_2SO_4 + Zn \longrightarrow ZnSO_4 + H_2$	(Cationic Single Replacement)
(t)	$HC_2H_3O_2 \rightleftharpoons H^+ + (C_2H_3O_2)^-$	(Ionization (Notice That It Is Reversible)

XV. Synthesis of the Module

This module entitled "Introductory Chemistry I" comprises of four (4) units namely: Matter and Measurements; Atomic Structure and Periodicity; Bonding, Molecules and Periodicity; and Chemical Reactions and Stoichiometry. Under "Matter and Measurements", the focus is on understanding the term matter and the corresponding distinguishing characteristics between its various forms. Emphasis is placed on identifying and explaining the differences between physical and chemical properties of materials and understanding the differences between physical and chemical changes. A discussion on the differences between matter and energy is also covered. The use of the International units of measurements (SI) is also given prominence in this unit.

Under the Unit "Atomic Structure and Periodicity", prominence is given to the various atomic theories in existence and the use of the modern atomic theory in explaining the structure and electronic configuration of atoms. In this unit, relevant terms such as atomic mass, mass number, and isotopes are defined and related calculations are carried out. A section of the unit also discusses and explains the thought process in the development of the periodic table; and explains the structure and properties of elements in groups and periods using the periodic table.

Unit III which deals with "Bonding, Molecules and Compounds" looks at the concept behind the arrangement of atoms of various elements to form new entities that themselves contain the individual atoms, linked together in some fashion. The nature of the interactions between the various atoms is considered in depth. Factors influencing the formation of ionic versus molecular compounds are also covered. Concepts behind the formation of different types of bonding (e.g., ionic versus covalent) are covered in detail. Also, the international system of nomenclature adopted in naming and writing the formulas of simple binary inorganic and organic compounds, cations and anions are considered. A section of the module deals with explaining and predicting the nature of a chemical bond, structure and geometry of simple molecules and ions. A description and application of Lewis theory in making Lewis structures of elements and binary compounds is also covered. In here, the procedure followed in drawing Lewis structures is demonstrated.

The last Unit of this module which deals with "Chemical Reactions and Stoichiometry", covers the mole concept and the interconversion between grams, moles, atoms and molecules. Processes involved in the determination of mass percent composition of a sample from experimental data is also covered. A considerable chunk of this unit is devoted to understanding the concept of basic stoichiometry and related calculations as well as the determination of percent composition of a compound from its formula, determination of empirical formula of a compound from its mass percent composition as well as writing of balanced chemical equations from descriptions of chemical changes.

XVI. Summative Evaluation

1. In each of the following:

   (a) Distinguish between density and specific gravity and state their common metric units.
   (b) Distinguish among the terms element, compound, and mixture.
   (c) What is the most important atomic property in Dalton's atomic theory?
   (d) What is the difference between an extensive and an intensive property? State three examples of each.
   (e) Define in your own words, the terms *precision* and *accuracy*.

2. State, in words, the following laws:

   (a) the law of conservation of mass.
   (b) the law of definite proportions.
   (c) the law of multiple proportions.

3. What properties distinguish solids from liquids? Solids from gases? And liquids from gases?

4. Write the orbital diagram for the electrons in an iron atom.

5. Specify metal (1) or nonmetal (2) for each of the following:

   (a) sulfur ____
   (b) chlorine ____
   (c) sodium ____
   (d) iron ____
   (e) carbon ____
   (f) silver ____

6. For each of the following (a – d) choose the correct one:

   (a) Element in Group VIIA, period 4:
   1) Br 2) Cl 3) Mn
   (b) Element in Group IIA, Period 3:
   1) beryllium 2) magnesium 3) boron
   (c) Metals in Group IVA:
   1) Ge, Sn, Pb 2) C, Si 3) C, Si, Ge, Sn

   (d) Nonmetals in Group VA
   1) As, Sb, Bi 2) N, P, As 3) N, P, As, Sb

7. State the number of protons for atoms of each of the following:

(a) Nitrogen
1) 5 protons   2) 7 protons   3) 14 protons
(b) Sulfur
1) 32 protons   2) 16 protons   3) 6 protons
(c) Barium
1) 137 protons   2) 81 protons   3) 56 protons

8. Naturally occurring carbon consists of three isotopes, ^{12}C, ^{13}C, and ^{14}C. State the number of protons, neutrons, and electrons in each of these carbon atoms.

	^{12}C	**^{13}C**	**^{14}C**
	6	**6**	**6**
#p	_______	_______	_______
#n	_______	_______	_______
#e	_______	_______	_______

8. An atom of zinc has a mass number of 65.

(a) What is the number of protons in the zinc atom?
(1) 30   (2) 35   (3) 65
(b) What is the number of neutrons in the zinc atom?
(1) 30   (2) 35   (3) 65
(c) What is the mass number of a zinc isotope with 37 neutrons?
(1) 37   (2) 65   (3) 67

9. Write the atomic symbols for atoms with the following:

(A) 8 p^+, 8 n, 8 e^- ___________

(B) 47 p^+, 60 n, 47 e^- ___________

(C) 17 p^+, 20 n, 17 e^- ___________

10. Using the periodic table, specify the atomic mass of each element (round off to the tenths place):

(A) calcium __________
(B) aluminum __________
(C) lead __________
(D) barium __________
(E) iron __________

11. For phosphorus, indicate if each configuration is (1) correct or (2) incorrect. Explain why or why not?

(A) 2, 2, 8, 5 1 or 2
(B) 2, 8, 3 1 or 2
(C) 2, 8, 5 1 or 2
(D) 2, 6, 7 1 or 2

12. Using the periodic table, write the complete electronic configuration for each of the following:

(A) Cl
(B) Sr
(C) I

13. We want to produce 2.75 mol of NH_3. How many moles of nitrogen would be required?

Use the following reaction equation to answer questions 13 – 16:

$2\ H_2 + O_2$ ---> $2\ H_2O$

14. How many moles of H_2O are produced when 5.00 moles of oxygen are used?

15. If 3.00 moles of H_2O are produced, how many moles of oxygen must be consumed?

16. How many moles of hydrogen gas must be used, given the data in problem two?

17. Suppose 4.00 grams of H_2 were used? How many grams of water would be produced? Here are two links intended to give you a hint as to how to solve this problem.

18. Write the molar ratios for PCl_3 to Cl_2 and PCl_3 to PCl_5 from the equation: $PCl_3 + Cl_2$ ---> PCl_5

19. Write the molar ratios for NH_3 to N_2 and H_2O to O_2. from the equation: 4 NH_3 + 3 O_2 ---> 2 N_2 + 6 H_2O

20. Write the molar ratios for PCl_3 to Cl_2 and PCl_3 to PCl_5 from the equation: PCl_3 + Cl_2 ---> PCl_5

21. Given this equation: 2 KI + $Pb(NO_3)_2$ --> PbI_2 + 2 KNO_3 calculate mass of PbI_2 produced by reacting of 30.0 g KI with excess $Pb(NO_3)_2$

22. How many grams of $AuCl_3$ can be made from 100.0 grams of chlorine by this reaction: 2 Au + 3 Cl_2 ---> 2 $AuCl_3$

23. How many grams of Na are required to react completely with 75.0 grams of chlorine using this reaction: 2 Na + Cl_2 ---> 2 NaCl

XVII. References

Ralph H. Petrucci and Williams S. Haward; **General Chemistry, 7th Edition,** Prentice – Hall International Inc., U.S.A, 1997.

Norbert T. Porile**; Modern University Chemistry,** Harcourt Brace Jovanovich, Inc., Nework, 1987.

Raymond Chang, **Chemistry, 4th Editio**n. McGraw-Hill Inc., New York, 1991, Oxford University Press, 2002.

Peter Atkins and Julio de Paula, **Atkin's Physical Chemistry**, 7th Edition.

Clyde R. Metz, **Schaum's Outline Series, Theory and Problems of Physical Chemistry**, McGraw-Hill Book Company, 1976.

Henry F. Holtzclaw and William R. Robinson, **General Chemistry 8th Edition**, D. C. Heath and Company, U.S.A., 1988.

Raymond Chang, **Chemistry 8th Edition**, McGraw-Hill, New York, 2005.

A.J. Walton; Three phases of Matter (2nd Edition), Clarendon Press, Oxford, 1983.

James E. Brady and Gerard E. Humiston, General Chemistry, Principles and Structure (3rd Edition), John Wiley & Sons, 1982.

B.H. Flowers and E. Mendoza; Properties of Matter, Wiley, London, 1970.

XVIII. Student Records

Name of the EXCEL file :

Continuous Assessment		Final Exam	Total
Assessment I. 25%	Assessment II. 25%	50%	100%

XIX. File Structure

Name of the module (WORD) file

Introductory Chemistry I.doc

Name of all other files (WORD, PDF, PPT, etc.) for the module

Atoms & Elements.pdf
Atoms and Elements 2.pdf
Atoms and Isotopes.pdf
Atomic Structure.htm
Atomic Theory I.htm
Atomic Theory II.htm
Bishop_Book_1_ebook.pdf

Bishop_Book_2_ebook.pdf
Bishop_Book_8_ebook.pdf
The Mole1.htm
Reactions.htm
Review of Elements, Compounds, and Mixtures.htm
Chemical Nomenclature.htm
Chemical Reactions.htm
ChemTeam Stoichiometry Molar Ratios.htm
ChemTeam Stoichiometry Mole-Mass Problems.htm
ChemTeam Stoichiometry Mole-Mole Problems.htm
Electron Arrangement and Periodic Law.pdf
Naming of Inorganic Compounds.htm
Naming Compounds, Ions, and Acids.htm
Nomen.htm
Notes Atoms & Ions.htm
Notes Molecules & Compounds.htm
Periodicity.htm
Physical and Chemical Changes.htm
Physical and Chemical Properties.htm
Reactions.htm
Reviews of Elements, Compounds, and Mixtures.htm
States of Matter.htm